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AN INTRODUCTION

TO

CHEMISTRY AND PHYSICS

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PREFACE

TO THE FIRST EDITION

THE assured position that Chemistry now holds in English education is shown by its inclusion as a compulsory or at least an optional subject in almost all examinations preliminary to entrance on the learned professions, and by the fact that many of the best schools now possess well-equipped laboratories and give regular instruction in the elements of both Theoretical and Practical Chemistry.

In directing the attention of our nation to the study of Chemistry as a means of scientific education, Sir Henry Roscoe has led the way. His text-books have given to generations of students their introduction to a knowledge of the science. His Primer of Chemistry and List of Experiments prepared for the Science and Art Department contain the subject-matter which every schoolboy has till recent times been expected to know, every examiner has asked for, and few text-books have dared to disregard. The output of admirable text-books on these lines has during recent years been so abundant, that the issue of another would be unjustifiable, if it differed from them merely in detail and in unimportant features.

But while the knowledge of the actual facts of Chemistry is worth much, there has been a growing feeling that the methods of gaining knowledge are often of more educational value than the knowledge itself, and that a lad who starts life with his mind merely stored with facts is not so well equipped as one who has

also *learnt how to learn*, not from books alone, but by direct contact with realities. In a school curriculum Science is chiefly valuable as a means of culture, and because there are faculties of the human intellect, namely, correct observation, sound reasoning from observation and experiment, and imagination in seeing into the invisible causes of visible phenomena, which can alone find their development in a carefully considered course in Physical Science. We are, therefore, even more anxious to insist on the methods by which knowledge has been obtained than on the actual facts of Chemistry. Sound methods of observation, moreover, remain valid, and may continually enlarge and correct our ideas in studying the further development of the science. At the same time the student must be sure of his facts before he begins to theorise. "Why is it," Charles II. is reported to have asked, "that when you put a dead fish into a bowl full of water it runs over, but if the fish is alive it does not run over?"

Some time ago Professor H. E. Armstrong, a leader in the movement which has sought to place the teaching of Chemistry upon what is regarded by many as a more rational basis, appended a detailed scheme for a school course on Chemistry to the Report upon Methods of Teaching Chemistry, which was presented to the British Association at Newcastle.[1]

Very recently, the Incorporated Association of Headmasters has issued a syllabus of a course of instruction in Elementary Science, including Physics and Chemistry, drawn up by a Committee on Science Teaching, of which Mr. C. M. Stuart, of Catford, was chairman, and Professor Armstrong a prominent member. To quote from this syllabus, "This course is intended for all boys and girls commencing the study of Science. It represents in the opinion of the Committee a suitable commencement for those who continue the subject, and indicates the manner in which it may be made of true educational value to those who do not pursue it further. . . . While the main object of the course should be to train students to solve simple

[1] See *Report of the British Association*, 1889. A subsequent paper elaborating certain features in this scheme may be found in *Nature*, vol. xliii. p. 593.

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problems by experiment—to work accurately, and with a clearly defined purpose, and to reason from observation—the instruction given should eventually lead them to comprehend the nature of air, water, 'fire,' earth and food."

This syllabus of the Association of Headmasters reached us after we had nearly completed our work on the present book, and it was a matter of satisfaction to us to find that almost the whole of the syllabus and the detailed experiments for the course on Elementary Chemistry were dealt with in our manuscript.

While the aim of the course described in the present book, and the object of the courses referred to above have alike been the cultivation of a spirit of inquiry and the prosecution of Science as among the best means of culture, there is, nevertheless, one difference between the lines on which they have been worked out.

The principle which has been our general guide in the selection and treatment of subject-matter may be expressed in words used in another connection by Professor H. A. Miers, in a recent inaugural lecture[1] at Oxford, in which he applied the saying of the Darwinists that the development of the individual is an epitome of the evolution of the race.

"The order," he said, "in which a subject can best be unfolded before a student's mind is very satisfactorily marked out by the historical development of the subject: a profitable course of teaching is suggested by the history of a science, and *the order in which problems have presented themselves to successive generations is the order in which they may be most naturally presented to the individual*." This is, we believe, particularly applicable to the study of Chemistry on the part of beginners. At the same time the progress of Science has often of course been like a paper-chase with many false scents, and it is not necessary for the logical application of the above principle to follow all these blind alleys.

A glance at the Table of Contents will indicate the course which we have been led to adopt in our attempt to work out an

1 *Nature*, vol. liv. p. 208

Introduction to the Study of Chemistry. We have referred at the outset to Alchemy and to some of the errors which were current till the seventeenth century, showing the readiness with which errors arose unless checked by well-devised experiment and careful measurement. In this way the student is led to see the *necessity of practical work and of exact measurement at a very early stage.* Measurements are then made of length, of mass, of the volume of liquids, of temperature, of density, of the pressure of the air, and of heat.

Then follows the practice of important chemical operations, such as solution, filtration, evaporation, and crystallisation, and the student's powers of observation and description are exercised by the examination of certain important raw materials, found native in the crust of the earth. He next learns how to prepare the most important acids and alkalies from his raw materials, after the manner of Geber, Glauber and other alchemists.

Having thus acquired in some measure quickness of eye in observation, accuracy of expression in description, delicacy of hand in manipulation, and a passing acquaintance with the apparatus and chemicals in the laboratory, the student is ready to set to work upon a series of researches upon the Action of Acids on Chalk, Fire and Air, the Rusting of Iron, the Discovery of Oxygen, the Action of Acids on Metals, the Composition of Water, and other subjects, in which he is led to work so far as may be along the lines which were successfully pursued by Black, by the Oxford chemists of the seventeenth century, by Priestley, Scheele, Lavoisier, Cavendish and by other famous investigators. Many of the experiments introduced have been suggested by their researches, and frequent quotations of interest have been made from their accounts of their own experiments.

The student will thus become acquainted with several Gases, and is therefore ready to investigate some of their principal properties. Researches follow which have a direct bearing on the law of Definite Proportions; for example, a complete research is made upon Chalk, following the lines of Black's classical investigation, and another upon the quantitative composition of Water. * * * * * *

A chapter on the Law of Definite Proportions closes the book.

We hold then that not only the *order*, but also the *method*, by which Chemistry can best be unfolded before a student's mind is suggested by its historical development. It is a necessary corollary that Chemistry cannot be learnt in the lecture-room alone, and that the *experiments should be for the most part quantitative in character*. Qualitative analysis has hitherto been largely used as the practical exercise in the science. It has some merits. We think, however, that practical chemistry should not consist, to such an extent as is usually the case, of "qualitative analysis," and that the right place for such a course is after and not before a simple quantitative course. While practical work should be introduced at a very early stage, the lecture-room has of course a necessary place, and the student should not start an experiment until he thoroughly understands what the object of the experiment is, what he means to do, and how he means to do it.

There are somewhat rigid limitations of time and circumstance which must be considered in selecting experiments suitable for a school laboratory. The experiments must not require more than two pairs of hands; the manipulation must not be so difficult that accurate results cannot be obtained; the apparatus must be cheap enough to allow of all the boys doing the same experiment at once, and the whole experiment must not take more than 1 or 1½ hours, including the time required for putting the apparatus away. These conditions exclude many important experiments; for instance, the determination of the composition of water by passing hydrogen over copper oxide may be done by the teacher as a lecture-table experiment, but it must generally be omitted by individual members of the class. Weighings should generally be conducted to the nearest centigram, because a boy's errors of manipulation are apt to exceed 1 centigram, and therefore milligram weighings will only give a fictitious accuracy and waste time. One per cent. accuracy is the standard that may be aimed at. We have dwelt at some length upon the details which are necessary to exactness in the

results of experiments. It is a great advantage to have several boys working at the same time at an experiment, for example, on the weight of a litre of air or on the percentage of water in a crystal; results should afterwards be compared and discussed. A great deal of time is saved by letting boys work in pairs, moreover, a spirit of co-operation is cultivated, and they learn at least as much as in doing the experiment alone, perhaps more, because they teach one another. * * * *

We would add that experience shows that a student's mathematics should be ready before he touches the apparatus of the laboratory. And this applies not only here. If any would-be chemists have not yet mastered the elements of arithmetic, decimals, the unitary method, percentages and proportional parts, we recommend them to close this book and go back to their ciphering. There can be no sound knowledge of Physics or of Chemistry without mathematical backbone. There is nothing more distracting to teacher and to student than to find that laboratory results cannot be worked out for want of adequate mathematical knowledge.

It may perhaps be objected that we have abandoned the principle we laid down as our guide in relegating Black's research on Chalk to so late a position. Our reply is that the aforesaid principle while invaluable as a general guide, must be applied with discrimination, and other considerations must not be lost sight of. For instance, it is desirable that a study of "every-day" phenomena of great interest, such as Fire and Air, should precede what is after all of a more special character. Moreover, the research on Chalk is the most complete in itself of any entered upon, and one which makes too great a demand upon the powers of hand and of mind to be introduced in its true chronological place immediately after the discovery of Fixed Air.

The value of this book would be little in practice, whatever its general merit, if it came straight from the writing desk, and had not been employed in actual class work. Not only have many of the chapters been worked through by elementary students in the laboratories of the Owens College, but we have been most fortunate in securing the co-operation of Hugh

Richardson, M.A., Science Master of Sedbergh School, and of S. H. Davies, M.Sc., Head of the Chemical Department of the Battersea Polytechnic, London. Using typed copies of our MSS., they have taught through almost the whole in their classes. We have amended our text wherever their trial of it showed alteration to be required, and we are convinced that the practicability and success of our scheme in actual class work has thus been secured and verified. Mr. Richardson also made at our request the first draft of the chapters on the Properties of Gases, and he has from the first helped us with his experience in the teaching of Science in schools.

We gladly thank also Mr. S. M. Walford, of the Hyde and Oldham Technical Schools; Mr. Foster, of the Nelson Technical School, and Mr. S. E. Brown, of the Friends' School, Ackworth, for testing some chapters in their own classes.

Our colleagues, Mr. G. J. Fowler, Dr. A. Harden, and Dr. W. A. Bone, have also given us kind assistance in certain details, while Mr. P. J. Hartog has helped us with criticism upon our chapter on the Law of Definite Proportions.[1]

While we have spared no trouble in a task which has been no light one, we cannot hope that we have avoided all error either in judgment or in points of detail, and suggestions and corrections offered by fellow teachers will be freely welcomed.

We may add that it is our desire to include in any subsequent edition chapters, which are already in hand, on the Discovery of the Metals and on Fuels and Food-Stuffs.

W. H. P.
B. L.

OWENS COLLEGE,
July, 1896.

[1] See *Nature*, vol. l, p. 440, and *British Association Report*, 1894, p. 518.

PREFACE

TO THE SECOND EDITION

We have been encouraged by the sympathetic reception which has been accorded to the first edition of our little book to submit it to a thorough revision.

The marked trend of elementary science teaching during recent years along the lines which we in common with many others have been advocating and practising is seen for instance in the present syllabus of Elementary Experimental Science of the Association of Headmasters, and in those of the Oxford and Cambridge Local Examinations, the Joint Scholarship Board, and the Scholarship Examinations of the London County Council.

These all show an increasing recognition of the vital necessity of keeping alive in both boys and girls that power of "making knowledge"[1] through their own experiences which, though possessed by every young child, is so commonly stifled and lost as years advance. To form a habit of inquiry, and to train the intelligence, and not the accumulation of facts, is now generally admitted to be the first aim of an elementary course in science. It is desired, in a word, *to teach boys and girls to think for themselves and to learn from their own observations.*

It has sometimes been alleged as an objection to the method which our book illustrates, that sooner or later some fact

1 Professor J. G. Macgregor, *Nature*, September, 1899.

must be assumed which a boy has not yet proved. That is no doubt the case, but his position is also that of the original investigator of to-day. John Dalton refused to accept the results obtained by others, and to his loss. The scientific method does not forbid such aid. Still, in an introductory course it is desirable that not many facts should be borrowed in the few typical researches pursued, in order that a very habit of research may be formed; afterwards the student may safely bring facts established by others to bear upon his own inquiries. Nor should it be forgotten that the scientific method includes deductive methods as well as inductive. Theories and working hypotheses inspire research, and without ideas progress stops. It is for lack of them that so few boys make any profitable use of a laboratory if left alone without the help of books or teachers. We have endeavoured to illustrate the importance of ideas by the use we have made, for instance, of Lavoisier's Theory of Combustion, when investigating the burning of Carbon and the explosion of Inflammable Air and Oxygen.

In this connection it may be added that the comparison of results and the discussion of them by the teacher is of the utmost importance. Instead of each boy repeating an experiment many times as the original investigator may do, he should be led to compare his own result with those obtained by his comrades, and to calculate the *mean* result. The comparison of results keeps alive a spirit of rivalry and a keen interest, and the mean result of quantitative experiments is usually surprisingly accurate, for example, the percentage of oxygen in the air or of carbonic acid in chalk can be determined with an error of less than one per cent. by means of the simplest apparatus.

There have been ominous warnings from Prof. G. H. Darwin[1] and from Prof. Karl Pearson[2] against the growing luxury of our laboratories and the costliness of apparatus; and this not because money is grudged for the cause of scientific education,

[1] Preface to his book on *The Tides*.

[2] *Lecture on Science and National Life*.

but because these material things cloud our eyes from seeing the really essential thing, the encouragement of the spirit of research. To this end we are content to work in rooms with simple fittings, and we have tried to avoid specialised apparatus, and to use combinations of rubber stoppers and flasks and tubes. This simplicity of apparatus results in great economy of expenditure and allows of a sufficient supply of balances and other essential instruments of general research. Moreover, as Prof. Ostwald has very recently pointed out, the simplest apparatus possible should be used that the attention may be fixed on the chief features of an experiment.

In this Second Edition the book is now divided into two volumes. The original volume covered two years' full work, and there is a distinct disadvantage in leaving a book long unfinished in a boy's hands. He begins to think he is stale of it. A new book is always a stimulus.

At the beginning of Volume I there are several chapters on Physical Measurement. Formal instruction, in our opinion, in the use of the rule, the balance and the burette is essential to progress. In some schools several hours are devoted to Science every week, and a more complete course in Physics than that we have introduced can be pursued. But in many schools other subjects will continue to make such demands upon the curriculum that the course of Physics, which should precede any study of chemistry, must not be too detailed. It is quite possible, of course, to spend a couple of terms or more upon measurements of the densities of all manner of materials, but the teacher will do well to recall the warning of Prof. J. J. Thomson, lest in teaching his boys to measure every physical quantity, he should deprive them of all wish to measure any of them. Grammar is not literature, nor is science measurement, though skill in the use of instruments of precision is essential. In collecting these physical chapters we have put together so much as seems to us necessary for progress in the chemical laboratory, but we rely on individual teachers to use them with discretion, selecting or postponing according to the circumstances of the class and the

school, and it must always be kept in mind that a boy, whilst learning to use his tools, will need to be inspired with a desire to employ them.

Volume II is purely chemical in character, and is intended to lead the schoolboy of 15 or 16 years of age along the paths of inquiry, which engaged the attention of the great chemists of the eighteenth century.

These two volumes contain the bulk of what was in our first edition. To this we have added some new chapters. The growing disposition to use Graphic Methods of representation and calculation, justify us in devoting a whole chapter to this subject, to which before we only referred incidentally. Many exercises are added to show the use of squared paper both for the records of results and as a means of arriving at new or unsuspected facts and relations. A new chapter on the Discovery of the Common Metals introduces simple blow-pipe work, minute smelting operations, which will give training in clean and neat manipulations and valuable practice in the examination of unfamiliar minerals. In another chapter on Fuels and Food Stuffs an examination of a variety of organic substances leads to a recognition of the fundamental importance of Carbon. Our treatment is in keeping with the educational maxim of taking familiar objects as texts, proceeding from the known to the unknown, and following the lines of the historical development of knowledge.

Three somewhat difficult chapters,[1] which are of less general interest, have been omitted, and are held over for a Third Volume, which we have in preparation. In our opinion they will find their proper setting in a Third Year's Course in which the Atomic Theory may arise for discussion.

The increasing number of Public Examination papers which recognise a change in the methods of teaching, and their great value both to boys and masters in directing attention to problems of interest, and in stimulating the very necessary

[1] Dalton's Law of Mixed Gases, The Combining Properties of Acids and Alkalies, and The Equivalent Weights of the Metals.

work of revision, have encouraged us to add, at the end of Volume II, some papers which have been set by Public Bodies. We think also that they are important as evidences of the growing harmony of view between teachers and examiners. Some private examination papers which have been set in schools using our book have been appended to Volume I.

It has been the fashion lately with some to decry any use of text-books. It is very noteworthy that young boys (æt. 12) learn very little from a chapter in a text-book, even after a lesson on the same subject, while it is equally remarkable how an older class (æt. 16) may realise, with fidelity of detail, experiments they have never seen, and make rapid progress with comparatively little help from a master. These two volumes are not intended to take the place of a teacher or of a boy's own note-book. The impressions which a boy's mind receives as he works in the laboratory must be expressed then and there in his own note-book, without the use of a text-book as a crib. But when a lengthy inquiry has been completed and he comes to write up his laboratory notes, some help and some additional information will often be needed, and then it is a part of a wise education to encourage the use of a text-book. We know that many teachers have found our book more useful to themselves than to their classes; an able boy can perhaps do without a text-book entirely, but his duller comrade often needs a helping hand in the revision of his work. It need scarcely be pointed out that there are experiments for which detailed and frequent instructions are necessary, and in such cases the necessary information may be obtained by a boy directly from a text-book, and its use will remove a burden from the teacher. And further, it is a great convenience to have mathematical examples in type. A few of them should be worked in class, and afterwards be set for revision without help. But enough of detail; every teacher must work in his own way to do his best work.

In conclusion, we have much pleasure in stating that Mr. Hugh Richardson, of the Friends' School, York, made for us

the first draft of the chapter on Graphic Representation, and we have throughout had the assistance of his experience. We gladly also thank Mr. S. H. Davies, late Head of the Chemical Department of the Battersea Polytechnic, for notes which we used in writing the new chapters on the Discovery of the Metals and on Fuels and Food Stuffs.

W. H. P.
B. L.

August, 1901.

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CHAPTER I

THE BIRTH OF CHEMISTRY

The Earliest Book on Chemistry.—Chemistry is one of the most modern of sciences, and one which has been placed upon a sure foundation only in the course of the last century. Yet the actual beginning and the dawning knowledge of Chemistry can be traced back to Egypt many centuries ago.

The oldest chemical record[1] of which we have any knowledge is a papyrus found at Thebes in the wrappings of the embalmed body of a goldsmith, who had lived about the time of the second or third century after Christ. From this note-book of the goldsmith, which is now in the library of the University of Leyden, we find that he was acquainted not only with gold and silver, but also with copper, lead, tin, zinc, and arsenic. Recipes are given for preparing imitation gold and silver, and for increasing the weight of gold and silver by the addition of inferior metals such as lead without altering the apparent character of the former.

The notes of this fraudulent workman, which have been preserved in perfect condition in an Egyptian mummy, and have escaped the accidents of fifteen hundred years, show that at that time the knowledge of the metals possessed by some of the Egyptians was very considerable. Not only could the author counterfeit true gold or silver, the "noble" metals as they were called, by mixing inferior or "base" metals, but

[1] See Berthelot, *Les Origines de l'Alchimie.*

tests are given by which the imitation could be detected and distinguished from the pure and genuine metal.

For instance, it is stated that gold should keep its colour when melted. If it becomes whitish it contains silver, if it blackens it contains lead, and if it becomes rough or hard when cold it contains copper or tin. So also silver should remain white and shining when fused. If it is blackened it must be falsified with lead, and if it is rendered hard and yellow, copper is present.

Alchemy: The Transmutation of Metals.—The imitation of the noble metals by the process of mixing the base metals in various proportions, as the writer of the papyrus described in the fourth century, was accepted in the next century as an actual change or *transmutation* of the base metals into the noble metals. Extraordinary as it may seem to us, for more than a thousand years such transmutation was believed in not only by the ignorant and unlearned but even by philosophers.

Indeed, it seems probable that the very word Chemistry is derived from the Greek *Chemeia* or χημεία, *transmutation*, since the transmutation of the metals was the chief object of the Egyptians.

About the year 640 the Arabians overran Egypt, and becoming acquainted there with Chemistry, they prefixed to the word for it the Arabic Article *al*, so that with them the science was spoken of as *Alchemy*. The Arabians spread through Northern Africa into Spain, and founded many Universities to which students flocked from all parts of Europe.

From the Egyptians the Arabian Alchemists had acquired the firm belief in the possibility of the change of one kind of material into another, and of the transmutation of the base into the noble metals. Their most earnest endeavours were given to the search for the *philosopher's stone*, a substance of such extraordinary virtue that according to Roger Bacon one part of it was believed to have the power to convert 1,000,000 parts of base metal into pure gold. The possibility of the existence of such a substance was accepted by the most learned: for instance, Van Helmont, a distinguished chemist, a physician and an honest man, recorded how in 1615 he obtained one quarter of a grain of the philosopher's stone and converted eight ounces of mercury into gold. Helvetius, again, was an opponent of the Alchemists until 1666, when he managed to obtain a little piece of the philosopher's stone, with which he said he

converted lead into a yellow metal, afterwards tested by the Master of the Mint, and pronounced to be gold.

In some cases we can trace the fraudulent means by which the apparent transmutation was effected. A certain Duke of Wurtemburg kept an Alchemist in his service, and following his directions, the Duke was accustomed to place the necessary ingredients in a crucible, and then, all having left the room, the door was locked, and the next morning, behold, pure gold was found in the crucible. It is recorded, however, that the gold was obtained, not by the miraculous agency of any philosopher's stone, but by the hand of the Alchemist's son, who had been concealed in a cupboard within the laboratory.

Remarkable as the belief in the transmutation of the metals appears to us, many facts and experiments may be mentioned which at first sight seem to confirm such a belief. The following experiments will illustrate this point :—

EXPT. 1. **To obtain Lead from Galena.**—Galena is a

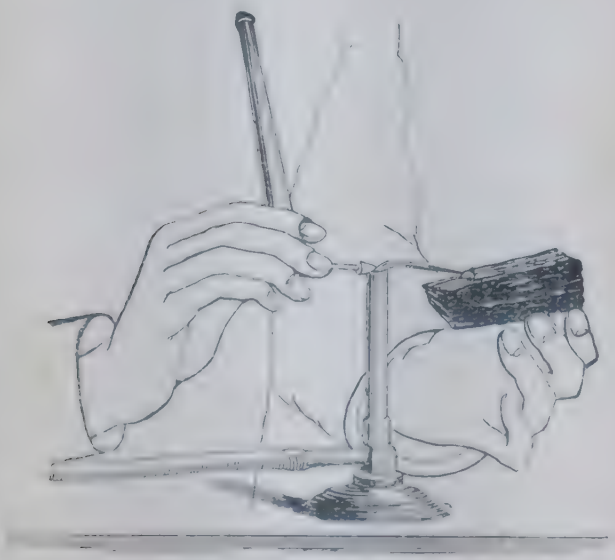


FIG. 1.

metallic-looking mineral. Grind a small portion to a fine powder in a mortar. Make a small cavity in a lump of charcoal, and place in it a little of the powder (as much as a small shot). Turn a Bunsen flame so as to form a small luminous flame. Place the nozzle of the mouth-blowpipe just within the

flame, and force a gentle current of air into it. Direct the tip of the flame upon the galena on the charcoal. Heat till white fumes are evolved, and a bright mirror-like surface of molten metal appears, which will perhaps gather into a bead. Allow it to solidify, and then transfer the bead to a mortar and crush it. It is flattened out into a little plate of metal. It is soft, marks paper, and has all the characteristics of lead.

The production of lead from galena was familiar to the Alchemists. Moreover, since by suitable treatment they found how to extract a small amount of silver from a large amount of lead, they supposed that an actual change of galena into lead and then of lead into silver was effected. It does not seem ever to have occurred to the Alchemists that the galena at the outset perhaps contained lead and silver. Their error, for it is now known to be such, might have been avoided if they had paid any attention to the weights of the substances employed and obtained in their experiments.

EXPT. 2. **To convert Zinc apparently into Copper.**—Pour some solution of blue vitriol (copper sulphate) into a test-tube till it is one-third full. Cut some strips of zinc narrow enough to enter the tube, drop in one strip, and notice the change of appearance. Boil, and observe that the zinc disappears, while a red and black deposit of copper is formed, as though the zinc had been transmuted into copper. Pour the solution into another test-tube, add another strip of zinc, and see whether any more copper is deposited on boiling.

EXPT. 3. **To convert Copper apparently into Silver.**—Clean a strip of copper thoroughly by means of sand-paper. Place it for a moment in a solution of sulphate of mercury.[1] On removing the copper and rubbing it with a flannel, it acquires a bright white silvery surface, as though the copper had been transmuted into silver. Heat the strip in a flame, and the bright silvery appearance is destroyed. The copper therefore has not at any rate been entirely changed into another metal, although, at first sight, such would appear to have been the case.

These experiments show that first impressions and ideas as to phenomena may be very far from the truth, and should in every

[1] Nitrate of mercury may be used in place of the sulphate, though the result is not quite so good.

case be severely tested and supported by further experiment before they are finally accepted.

Conclusions must be tested by Experiment and by Measurement.—Another illustration of the way in which ideas may change is found in the opinions which have been held as to whether dead matter of itself can ever become alive.

The Roman poet Vergil[1] records the notion as being widely prevalent in Egypt, Persia, and the East, that bees could be generated from the putrid carcases of oxen. He specially recommends a two-year-old bullock for the purpose, and prescribes the mode of killing it by suffocation and beating, then the pounding and mashing of it, hide and all, and laying it on thyme and cassia in a small shady chamber, with access of air at fit intervals, and so forth. Vergil does not tell us that he had seen it done, nor does he hint that the occasional selection by a swarm of bees of the hollow of a beast's skeleton as a convenient home may have originated the idea on which he enlarges with so much minuteness.

Our forefathers certainly believed that dead matter could beget life, and would have instanced vinegar, which breeds vinegar-eels. Yet this great question of spontaneous generation has since been answered by experiments, which show that all life is derived from pre-existing life—"*omne vivum e vivo*."

The laborious and life-long studies of the Alchemists did much to increase the knowledge of chemical operations, of minerals and metals, and of chemical agents, but the true method of scientific investigation was not theirs. They were too ready to be satisfied with a shallow explanation of what they saw. They were not accustomed to weigh the materials with which they began an experiment, nor to determine the weight of anything left at the end. In fact, measurement was neglected by them. The result was that they were led to conclusions which were to a large extent erroneous. Nature does not reveal her secrets to the careless, nor to impatient inquirers. And until the chemist made the balance his constant servant, and was willing to test all his conclusions by rigorous experiment, there was and could be no sure foundation for the theories and laws which he endeavoured to construct.

[1] Vergil, *Georgics IV.*, 281–314.

CHAPTER II

THE METRIC SYSTEM. THE MEASUREMENT OF LENGTH AND OF AREA

Standards of Measurement.—It has already been insisted upon that measurement is the first and most important step in the progress of any science, and it will be observed that in the statement of any measurement there is the mention firstly of a number, and secondly of a thing of the same kind as the quantity to be measured, which is referred to as a *standard* or *unit*. For instance, if it is desired to state the sun's distance from the earth, the distance of the moon from the earth may be selected as a standard or unit, and it may be stated that the sun's distance is 400 moon's distances. Similarly, the moon's distance may be stated in terms of the earth's radius, as about 60 earth's radii.

But whilst such statements convey definite meanings to the mind, since in either case the magnitude of the unit chosen is comparable with the quantity to be measured, yet it would be very inconvenient in scientific work to state quantities in terms of standards not universally recognised, for it would be extremely troublesome to compare the statement of one man with that of another if they employed different standards. It is on this account that a statement of quantities in yards, gallons, or pounds is almost meaningless to a Frenchman or German. On the other hand, the "Metric System" of measurement, to which reference is made below, is understood by scientific men of all nationalities, and indeed it is almost exclusively employed in scientific investigations.

There is another objection to the English system for purposes of measurement in the fact that there is no simple relation between the different standards of length, volume, and weight. There is consequently much trouble in making calculations with quantities expressed by means of them. The Metric System, on the other hand, is a *decimal* system, and the different standards are related in a simple way to one another.

We shall now refer to the standards of Length, Area, Volume, and Weight in the Metric System, which are used in chemical investigations, and explain the nomenclature used in connection with this system.

1. STANDARD OF LENGTH. — The **Metre**[1] is the length at 0° C. of a certain bar of platinum kept in Paris.

In 1790, immediately after the French Revolution, a commission of *savants* nominated by the French Academy was appointed to prepare a new system of units. This commission decided in favour of a decimal system, with the unit of length, from which all other units should be derived, itself connected with the size of the earth. This happy idea put the unit beyond reach of change from temperature or from other causes.

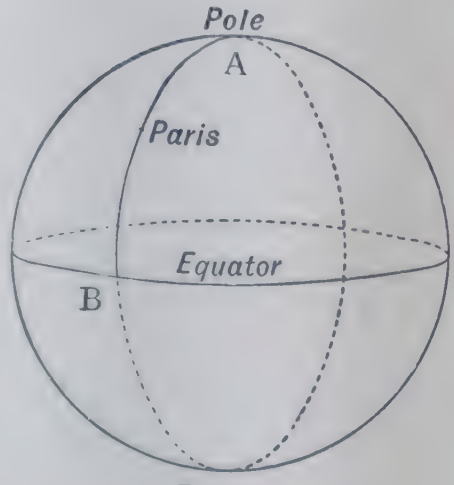


FIG. 2.

The commission added to itself Lavoisier and other eminent men, and had a meridian arc from Dunkirk to Barcelona carefully measured, as a basis for calculating the size of the earth. The metre was intended to be $\frac{1}{10{,}000{,}000}$ of the meridian quadrant of the earth's circumference drawn through Paris (Fig. 2). This would make the circumference of the earth 40,000,000 metres. A *platinum bar* of the required size was constructed. Since then a small error has been found in the calculations, so that this platinum metre is not exactly what it was intended to be. But this

[1] Greek, *metron*, a measure.

metre bar is still taken as the ultimate standard of appeal, and careful copies have been made and distributed among the European nations.

Examine a Metre Rule. Notice firstly that there are a great number of short lines upon it. Moreover, groups of these are marked off by longer lines, which are numbered, and there are seen to be 10 divisions in each group. There are 100 of these groups, and therefore there are 1000 of the smallest divisions in the whole metre.

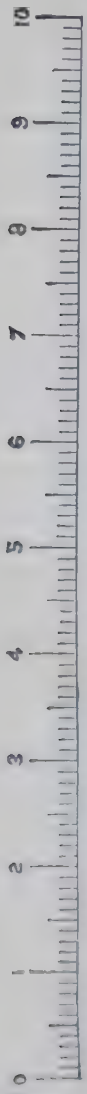


FIG. 3.

The width of one of the smallest divisions is called **1 millimetre** (Latin, *mille*, a thousand).

The width of a group of 10 millimetres is contained 100 times in the metre; hence this length is called **1 centimetre** (*centum*, a hundred).

The width of ten centimetres is contained 10 times in the metre; hence this length is named **1 decimetre** (*decem*, ten). Fig. 3 represents 1 decimetre; the long lines mark off centimetres, and the smallest divisions mark off millimetres.

Other names with Greek prefixes are given to lengths 10, 100, and 1000 times the metre. The most important of these is **1 kilometre**, which equals 1000 metres.

TO BE LEARNT BY HEART

1 millimetre equals	$\frac{1}{1000}$	of a metre; it is denoted by				mm.
1 centimetre	„	$\frac{1}{100}$	„	„	„	cm.
1 decimetre	„	$\frac{1}{10}$	„	„	„	dm.
1 metre				„	„	m.
1 dekametre	„	10	metres	„	„	Dm.
1 hectometre	„	100	„	„	„	Hm.
1 kilometre	„	1000	„	„	„	Km.

The student should try to remember roughly these metric measures by reference to real objects of known size. Thus:—

Kilometre equals distance covered by 10 minutes' quick walking.
Hectometre „ „ „ length of football ground.

Dekametre	equals	distance	covered by	half length of cricket pitch.
Metre	„	„	„	width of door.
Decimetre	„	„	„	width of hand.
Centimetre	„	„	„	width of finger-nail or dice.
Millimetre	„	„	„	nib of J pen, or thickness of sixpence.

THE MEASUREMENT OF LENGTH

Obtain a small box-wood rule[1] marked with a centimetre scale along one edge, and an inch scale divided into tenths along the other. Obtain a sharp-pointed hard pencil.

EXERCISES[2]

1. Draw a line 3 inches long. Measure its length in millimetres, to the nearest whole millimetre.

2. Draw a line 2·5 inches long. Measure its length to the nearest millimetre.

3. Draw a line 4·3 inches long. Measure its length in centimetres and millimetres. Then state the length in centimetres and decimal.

4. Measure the distance between the bottom and top lines in your note-book in centimetres and decimal.[3]

5. Measure the length of your note-book in decimetres, centimetres, and millimetres. Then state the length in decimetres and decimal.

6. Measure the breadth of your bench in decimetres and decimal.

7. Measure the breadth of the laboratory in metres and decimal.

The Estimation of Hundredths of an Inch and of a Centimetre.—Observe an inch scale divided into tenths. It is easy to make measurements by means of it to the nearest $\frac{1}{10}$ inch. But we can be more exact. We may imagine each $\frac{1}{10}$ inch to be divided into 10 smaller but invisible divisions of $\frac{1}{100}$ inch each.

1. They are provided by J. Rayborne and Sons, Whitmore Street, Birmingham, or by J. Thornton, 109, Deansgate, Manchester.

2. The eye should in each case be placed immediately above that part of the scale which is being read.

3. If a length is an exact number of cm., *e.g.* 13 cm., it should be stated as 13·0 cm.

Fig. 4 represents the edge of an inch scale.

AB = 0·7 inch to the nearest $\frac{1}{10}$ inch
= 0·7+ inch, *i.e.* between 0·7 and 0·8 inch
= 0·73 inch, estimating hundredths of an inch.

Similarly, a centimetre scale divided into millimetres can be read to the hundredth of a centimetre.

Exercises

Note.—In the following exercises read to 2 places of decimals.

1. Measure the distance between the bottom and top lines in your note-book in inches.

2. Number the lines in your note-book 1, 2, 3, 4, &c. Measure the distance between the tenth line and the first line in inches.

Fig. 4.

3. Measure the distance between the seventh line and the first line in inches.

4. Measure the distance of the first line from each of the following lines in inches. Tabulate your results.

5. Measure the distance between the tenth line and the first line in centimetres.

6. Measure the distance between the seventh line and the first line in centimetres.

7. Place a 10-to-the-inch scale against a scale divided into centimetres and millimetres. Read off in inches and decimals the length of 1 cm., 2 cm., 3 cm., ... 10 cm., estimating to the nearest hundredth of an inch. For instance,

1 cm. = ·4 inch.
2 cm. = ·79 ,,

8. Place two scales together as before. Read off in millimetres (and by estimation tenths of a millimetre) the lengths of 1, 2, . . . 6 ins.

9. Measure the diameter of a penny in centimetres.

10. Take a strip of paper about 15 cm. long. Wrap it carefully round the edge of a penny until it overlaps, and then prick a small hole through both thicknesses of paper with a pin. Measure in centimetres the distance between the two pin-pricks on the paper. This

length is equal to the circumference of the penny. Then divide the length of the circumference by that of the diameter. State thus :—

$$\begin{aligned} \text{Circumference} &= \quad \text{cm.} \\ \text{Diameter} &= \quad \text{cm.} \\ \therefore \frac{\text{Circumference}}{\text{Diameter}} &= \frac{\quad}{\quad} \text{ cm.} \\ \text{,,} &= \quad \text{cm.} \end{aligned}$$

11. Measure in the same way the circumference of

(*a*) a copper cylinder,
(*b*) a boxwood ball.

12. Repeat No. 11, but wrap the paper *three* times round the ball and cylinder.

13. Draw a circle of 5 cm. radius. Find the length of its circumference in cm. by means of a piece of cotton. Tie a knot near one end of the cotton. Mark any point on the circumference, place the knot upon it, and with finger and thumb gradually trace the cotton round to the mark again. Then pick up the cotton, and lay it straight along a metre scale. Lastly divide the length of the circumference by that of the diameter. State as in No. 10.

14. Measure the length of a road on an Ordnance Map by means of a piece of cotton, and then by reference to the scale of the map calculate the length of the road in miles.

Examples I

Metric Prefixes. Length

1. How many centimetres make $\frac{1}{2}$ metre?
2. How many millimetres make 1 decimetre?
3. How many metres make 10 kilometres?
4. How many millimetres make 1 kilometre?
5. Express 1 centimetre as a decimal of 1 decimetre.
6. Express 1 millimetre in kilometres.
7. Express 45·6 cm. in dm.
8. Bring 2·34 km. to cm.
9. Subtract 1 mm. from 1 cm.; answer in cm.
10. How many cm. are there in 123·456 metres?
11. The diameter of a halfpenny is 25 mm. How many halfpennies in a row would stretch 1 metre?
12. How many metre strides shall I take in running 5 kilometres?
13. If a sixpence is 1 mm. thick, find the height in cm. of a pile of sixpences worth £1.

14. If a cricket pitch is 2 dekametres long, how many kilometres does a boy run in making a century?

15. A fishing-rod has 3 joints and a total length of 3·6 metres. Find the average length of each joint in cm.

16. How many hours will it take to walk 24 kilometres at the rate of 1 hectometre per minute?

17. How many pins, each 25 mm. long, can be made from 1 kilometre of wire?

18. Four pieces of 1 decimetre each are cut away from a metre of glass tubing. Into how many lengths of 15 cm. long can the remainder be divided?

2. STANDARD OF AREA.—Draw a square ABCD measuring 1 decimetre along each side (Fig. 5). This is called **1 square decimetre.** Mark off centimetres along each side. Through the points of division rule lines parallel to the sides of the square.

The square decimetre is then divided into smaller squares, each **1 sq. centimetre.** How many sq. cm. does it contain? There are 10 rows, each containing 10 sq. cm., or 100 sq. cm. altogether. The sq. cm. AEFG has been similarly divided into sq. mm. Thus,

1 sq. dm. = 100 sq. cm.
1 sq. cm. = 100 sq. mm.

Observe that the area 1 sq. cm. is roughly represented by the area of a finger-nail.

A square millimetre, a square centimetre, a square decimetre, a square metre, or a square kilometre may in different cases be used as the unit of area. Thus the area of a country may be suitably expressed in square kilometres, the area of a playground in square metres, and the area of the wing of a bee in square millimetres.

The Measurement of Rectangular Areas.—Obtain some paper divided by straight lines into sq. mm.

(1) With a sharp-pointed pencil draw upon it a sq. centimetre. It contains 100 sq. mm.

(2) Draw upon it a rectangle 8 cm. long and 34 mm. broad. Find its area by *counting* the sq. mm. which it contains. It is

FIG. 5.

seen to contain 24 blocks of 100 sq. mm., and also a long strip containing 320 sq. mm. Thus :

$$\text{Area} = 2400 + 320 \text{ sq. mm.}$$
$$= 2720 \text{ sq. mm.}$$

Observe that we can also calculate the area by considering that it contains 34 rows of sq. millimetres, with 80 sq. millimetres in each row, or

$$\text{Area} = 34 \times 80 \text{ sq. mm.}$$
$$= 2720 \text{ sq. mm.}$$

(3) Draw a rectangle 6·9 cm. long and 5·3 cm. broad. What is its area?

$$\text{Area} = 69 \times 53 \text{ sq. mm.}$$
$$= 3657 \text{ sq. mm.}$$
$$= 36{\cdot}57 \text{ sq. cm.}$$

These examples will serve to show that in the case of any rectangular surface

Area = length × breadth.

Care must be taken to express both length and breadth in terms of the *same* unit.

Exercises

1. Obtain some paper divided by straight lines into square millimetres. Draw upon it a rectangle 5·5 cm. long and 24 mm. broad. Find its area

(*a*) in sq. mm.
(*b*) in sq. cm.

2. Draw a square inch upon the same paper. How many sq. mm. does it contain?

3. Obtain some paper divided by straight lines into square inches and hundredths of a square inch. Draw upon it a rectangle 2·2 inches long and 0·8 inch broad. Find its area

(*a*) in hundredths of a square inch
(*b*) in square inches.

4. On the same paper draw a rectangle 2·8 inches long and 1·7 inches broad. Find its area in square inches.

5. Open your note-book at the middle. Measure the length and breadth of the inside sheet in cm. (to 1 place of decimals). Calculate its area in sq. cm. and decimal.

6. Find the area of (*a*) a sheet of paper, (*b*) a postcard, in square inches and decimal.

7. Find the area of a visiting card in sq. cm. and decimal.

8. Trace the outline of a nasturtium leaf upon some paper divided into sq. mm. Calculate its approximate area in sq. mm.

Examples II

Metric Square Measure. Area

1. How many sq. dm. are there in 1 square metre?
2. How many sq. cm. are there in 1 square metre?
3. Multiply 3 dm. by 2 cm.: answer in sq. cm.

4. Multiply 80 mm. by 30 mm ; answer in sq. cm.

5. If 1 inch equals 25 mm., how many sq. mm. equal 1 square inch?

6. Find in sq. cm. the area of a postcard 12·5 cm. long and 80 mm. broad.

7. How many pieces of zinc each 25 cm. by 0·8 cm. can be cut from a sheet 16 dm. by 8 cm.?

8. How many postage stamps each 24 mm. by 20 mm. are required to paper a wall 4 metres by 3 metres?

9. If the pressure of the air is $2\frac{1}{2}$ lbs. per sq. cm., what is that in cwts. per square metre?

10. If 1 dm. equals 4 in. long, express 1 square metre in square yards.

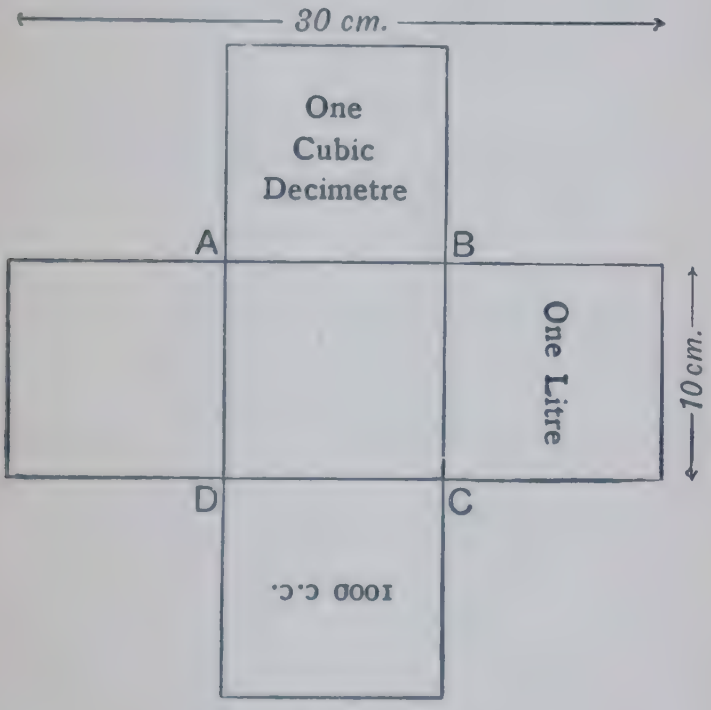


FIG. 6.

3. **STANDARD OF CUBIC CONTENTS OR VOLUME.**—The unit of volume most used in laboratory work is the volume of a cube each edge of which measures 1 cm. long. This is called **1 cubic centimetre** : it is denoted by **1 c.c.**

EXERCISE. **To cut out a Cubic Centimetre.**—Cut out of soap a cube, each edge of which shall measure 1 cm

Make a mental note of its size. It is about that of a die. The capacity of a small thimble is also about that of 1 c.c.

EXERCISE. **To make a Cubic Decimetre.**—Cut out a stiff cardboard cross, 30 cm. by 10 cm. (Fig. 6.) Print inscriptions on the faces as indicated in the figure. Cut the card half-way through along each of the lines AB, BC, CD, DA. Bend back the four squares and stick their edges together with gummed paper both inside and outside. Let the gum dry. Then warm the hollow cube, and paint it over with melted paraffin.

Similarly make out of cardboard a cubic centimetre.

How many cubic centimetres could be packed into the cubic decimetre? Imagine first that the bottom of the cubic

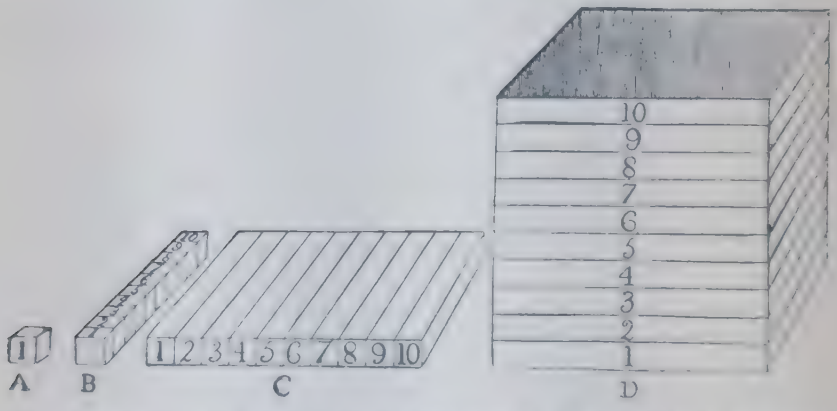


FIG. 7.—Ten cubes like *A* would make a row like *B*. Ten rows like *B* would make a layer like *C*, and ten such layers would go into the box *D*.

decimetre were paved with cubic centimetres, how many would cover the bottom? Clearly 10 rows of cubic centimetres with 10 cubic centimetres in each row, or 100 altogether, would be necessary. And ten of these layers piled one on another would be required to fill the cubic decimetre (see Fig. 7). Thus $10 \times 10 \times 10$ or 1000 cubic centimetres equal 1 cubic decimetre. The capacity of 1 cubic decimetre is called **1 litre.** On the Continent milk and oil and other liquids are measured in litres.

Take a large flask which will contain just 1 litre of a liquid, fill it with water, and find how many times you can fill a tumbler from it. It can be filled four times, so the capacity of a tumbler is about $\frac{1}{4}$ of a litre.

Fill a litre flask with water from a pint measure. It can contain about $1\frac{3}{4}$ pints.

The multiples and decimals of a litre are named like those of the metre; for instance, a *decilitre* means $\frac{1}{10}$ of a litre, and remembering that 1 litre = 1 cubic decimetre = 1000 c.c., we may construct the following table :—

1 litre		= 1000 c.c.
1 decilitre	= $\frac{1}{10}$ litre	= 100 c.c.
1 centilitre	= $\frac{1}{100}$ „	= 10 c.c.
1 millilitre	= $\frac{1}{1000}$ „	= 1 c.c.

We have thus two names for $\frac{1}{1000}$ litre; it may be called either 1 millilitre or 1 cubic centimetre. The latter is more often used.

It has already been shown that the volume or capacity of any rectangular solid is obtained by multiplying the area of its base by its height. Thus, if a biscuit-box is 5 cm. broad, 8 cm. long, and 6 cm. high, its capacity is $5 \times 8 \times 6$ or 240 cubic centimetres. Of course, before multiplying breadth × length × height, the numbers expressing the lengths of these must all be in terms of the *same* unit.

Exercises

Find the volume of the rectangular block of wood provided :

(*a*) in cubic inches,
(*b*) in cubic centimetres.

Measure inches to 2 places of decimals and centimetres to 1 place of decimals.

State thus :—

Length = inches
Breadth = „
Height = „
∴ Volume = × × cubic inches
= cubic inches

EXAMPLES III[1]

Metric Cubic and Capacity Measure

1. How many c.c. are there in 1 cu. dm.?
2. A cube measures 30 cm. each way; what is its volume in c.c.?
3. How many cu. mm. are there in 1 cubic inch, if 25 mm. equal 1 inch long?
4. How many cu. mm. are there in 1 c.c.?
5. How many c.c. are there in one cubic metre?
6. How many c.c. are there in 3 litres?
7. How many c.c. are there in 2 decilitres?
8. Express 1234 c.c. in litres.
9. How many c.c. are there in 56·78 litres?
10. Express in c.c. 1 litre + 2 decilitres.
11. The lead of a pencil is 2 mm. broad, 2 mm. thick, and 150 mm. long, what is its volume in cu. mm.?
12. Find in c.c. the contents of a biscuit-box 15 cm. high, 12 cm. broad, and 240 mm. long.
13. Find in litres the contents of a box, 8 cm. deep, 25 cm. long, and 10 cm. broad.
14. How many times can a dekalitre bucket be filled from a trough 2 metres long, 50 dm. broad, and 40 cm. deep.
15. 600 cu. ft. of air each is allowed in a dormitory for 20 people. How many litres is that for all? (1 ft. long = 30 cm.)
16. How many days will a hectolitre cask of water last for 7 people, allowing each 3 pints per day (1 litre = 1¾ pints).

4. STANDARD OF WEIGHT.—The unit or standard of weight in the Metric System is the **gram**. One gram (gm.) is defined as the weight of 1 cubic centimetre of pure water.[2]

A penny weighs about 10 grams.

It has already been stated that 1 gram is the weight of 1 c.c. of water. There is thus a very simple relation between the measures of the volumes and weights of given quantities of water. For instance, 1000 c.c. of water should weigh 1000 grams. Test this as follows:—Take a litre (1000 c.c.) flask, and counterpoise it on a strong balance; then add 1000 grams to

[1] *To the Teacher.*—All the questions in this set and in the succeeding sets of Examples in this chapter should not be worked on first studying the subject. Some of the questions should be set in the course of the study of the next few chapters.

[2] Strictly speaking at 4° C., the temperature of maximum density of water.

the counterpoise, and fill the flask with water; the balance is still maintained.

There is no such simple relation between the English measures of volume and weight.

Multiples and decimals of 1 gram are named *as before*. There is therefore no difficulty in learning the following table, whereas Avoirdupois Weight is not made at all easier by having previously learnt Long Measure :—

1 kilogram	=	1000 gm.
1 hectogram	=	100 gm.
1 dekagram	=	10 gm.
1 gram		
1 decigram	=	·1 gm.
1 centigram	=	·01 gm.
1 milligram	=	·001 gm.

Examples IV

Metric System. Weight

1. How many decigrams are there in 1 gram?
2. How many milligrams are there in 1 centigram
3. Express 7891 grams in kilograms.
4. What decimal of a gram are 2 decigrams?
5. What decimal of a gram are 50 milligrams?
6. Express 6·54 kilograms in grams.
7. What decimal of a gram are 100 milligrams?
8. Express 3 centigrams + 4 milligrams in milligrams.
9. Express 2 decigrams + 5 centigrams in milligrams.
10. Add 3 gm. + 2 dgm. + 50 mgm. Express in grams.
11. Add 25 gm. + 200 mgm. + 2 mgm. Express in grams.
12. Add 1 Kgm. + 2 Dgm. + 3 gm. + 4 dgm.
13. A flask weighs 20 gm. when empty, and 45 gm. when full of water. What is its capacity?
14. A flask which will hold 80 c.c. of water weighs 107 grams when full of water. How many grams will it weigh when empty?
15. A litre flask weighs 150 gm. when empty. How many grams will it weigh when full of water?
16. A biscuit tin measures 16 cm. by 25 cm. by 20 cm. How many grams of water will it hold?
17. Mercury is 13½ times as heavy as water. How many grams of mercury will fill a decilitre flask?

18. Sulphuric acid is 1·8 times as heavy as water. How many c.c. of sulphuric acid will weigh 99 grams?

19. If 5 cm. of wire weigh 25 milligrams, how many grams will 10 metres of the wire weigh?

20. If 1 square centimetre of metallic foil weighs 1 centigram, how many grams will 1 square decimetre weigh?

The Relations between English and Metric Units:—

Some rough equivalents are worth remembering:—

1 foot	equals	30 cm.
1 inch	,,	25 mm.
1 decimetre	,,	4 in.
1 kilometre	,,	$\frac{5}{8}$ mile.
1 ounce	,,	28 grams.
1 litre	,,	$1\frac{3}{4}$ pints.
1 kilogram	,,	$2\frac{1}{6}$ lbs.

More exact numbers are useful for reference:—

1 metre	equals	39·37 inches.
1 foot	,,	30·5 cm.
1 gram	,,	15·43 grains.
1 grain	,,	·06489 grams.
1 kilogram	,,	2·20 lbs.
1 litre	,,	1·76 pints.
1 litre	,,	61·03 cu. ins.
1 cu. in.	,,	16·38 c.c.

Note.—If 1 gram equals 15·43 grains, then 1 gram equals $\frac{1}{15{\cdot}43}$ gram, or ·06489 gram.

The numbers 15·43 and ·06489 are called *reciprocals*.

Some Disadvantages of the Metric System.—The advantages of the Metric System over any existing system for the purposes of scientific investigation are great. It is understood by men of all nationalities, and the decimal relations which exist between the units are of great convenience in calculations.

It is not, however, urged here that the Metric System is the best for all the ordinary transactions of life. Indeed, as Herbert

Spencer has recently pointed out, there is some ground for the belief that a decimal system is not the most convenient for the purposes of petty trade. Quoting from a letter which appeared in the *Times*, he gives the following instance: "In England, if an old village woman buys a measured quantity involving fractions of an article at a price reckoned, say, in shillings, pence, and halfpence, the shopman or girl never seems to find any difficulty in working quickly in his or her head the sum which has to be paid, and the old woman manages to see that she is charged correctly. But, in France, with the decimal system, it seems always to be necessary to work the simplest sum out on a piece of paper. In fact, it is not difficult to see that it is pretty easy to work out mentally the value of, say, a yard and three-quarters of riband at $3\frac{1}{2}d.$ a yard, while it is very difficult to work out mentally the value of, say, 1 m. 75 centimetres of riband at 35 centimes a metre."

Moreover, in different trades different units are found to be convenient, and though the Government of France enforced as long ago as 1839 the observance of the metric system, it is remarkable that its use is even to-day not general in France. "Precious stones are to-day bought and sold in carats; firewood in cordes; milk in pintes; gravel in toises; grain, potatoes, and charcoal in boisseaux; sugar and tea among the poor people is dealt with in livres, demilivres, &c."

Examples V

Metric System. Miscellaneous

(*a*)

1. How many pieces of glass rod, each 5 cm. long, can be cut from a piece 1 metre long?

2. How many times can 10 c.c. of water be obtained from a half-litre flask?

3. If 1 metre of wire weighs 4 decigrams, how many milligrams will 5 cm. weigh?

4. A flask weighs 20 grams when empty, and 120 grams when full of water. What will the flask weigh when full of mercury, of which 1 c.c. weighs 13·5 grams?

5. A train is travelling 100 kilometres per hour. What is that in centimetres per second?

6. If a metal decigram weight measures 1 sq. cm., find in grams the weight of a piece of metal of the same thickness, measuring 4 mm. by $2\frac{1}{2}$ mm.

(*b*)

7. Add, expressing the answer in metres :—
1 Km. + 2 mm. + 3 cm. + 4 Hm. + 5 Dm. + 6 dm.

8. Find in sq. cm. the area of a post-card measuring 1·25 dm. by 80 mm.

9. Find in cu. dm. the contents of a box measuring 150 mm. high, 12 cm. broad, and 2·4 dm. long.

10. How many grams of soda are contained in 10 c.c. of a solution, 1 litre of which contains $\frac{1}{25}$ of a kilogram of soda?

11. If 1 metre of wire weighs 120 centigrams, what will be the length in centimetres of 60 milligrams of the same wire?

12. Find to 2 places of decimals the number of centimetres in 1 foot long, assuming that 1 metre = 39·37 inches.

(*c*)

13. Add together $\frac{1}{2}$ litre, 20 c.c., 200 millilitres, and 12 decilitres, and express the sum in c.c.

14. If a bath is 17·1 metres long, 5·1 metres wide, and 1·5 metres deep, how many times would a bucket, containing 8·1 litres, have to be filled before the bath was *half* empty?

15. If I deal out to each of 40 boys 3 pieces of wire, each piece being $\frac{1}{8}$ metre long, and if a metre of wire weighs 4 decigrams, what weight of wire have I dealt out?

16. If the diameter of a sixpenny piece is 2 cm. long, and it is 1 mm. thick, calculate what would be the value of a hectometre of them placed in a *row*.

17. Calculate how high the same number of coins would reach, expressed in metres, if they were placed in a *pile*.

18. If a bicyclist rides 10·5 miles in 42·5 minutes, and if 8 kilometres = 5 miles, how many kilometres an hour does he travel on the average?

(*d*)

19. Add 2 kilograms, 400 grams, 200 milligrams, 5 decigrams, 1 centigram. Answer in grams.

20. If a square centimetre of platinum foil weighs 1 centigram, how many grams will a square metre weigh?

21. How many times can a decilitre flask be filled from a pneumatic trough 10 cm. broad, 8 cm. deep, and 25 cm. long?

22. A solution contains 1 ounce of salt in 1 pint of water. How many grams of salt per litre of water would make a solution of the same strength?

23. If a 100 c.c. flask weighs 120 grams when full of water, what will it weigh when full of mercury, which is $13\frac{1}{2}$ times as heavy as water?

24. If 30 cm. = 1 linear foot, and 1 cubic foot of water = 1000 ounces, find the weight of 1 litre of water in ounces.

CHAPTER III

THE MEASUREMENT OF MASS—THE BALANCE

THE mass or quantity of matter in a substance is measured by comparing it with standard masses. These standard masses are usually called "weights." Hence, the measurement of mass is commonly spoken of as "weighing." The measurement of mass is usually carried out with an apparatus called a *balance*.

EXERCISE.—**To make a Simple Balance.**[1]—Take a box-wood scale, and drill a hole near one edge at the division

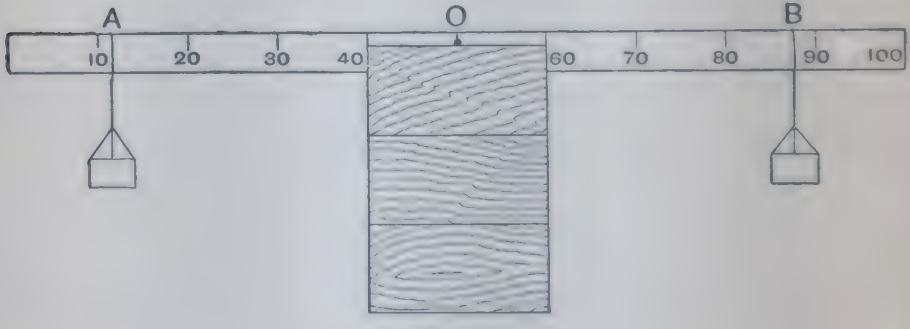


FIG. 8.

marked 50, and fix a steel knitting-needle tightly through the hole. Place two piles of wooden blocks, with their faces parallel to one another, and about 1 inch apart. Let the knitting-needle rest on the upper surfaces of the blocks with the scale between them. Does the scale balance horizontally? If not, wind some tin foil round one end until it does. Take two pill-boxes without lids and arrange to suspend them by thread from

[1] Adapted from *Weighing and Measuring*, by F. Edser.

the scale. Hang them at equal distances from O, the middle of the scale, and on opposite sides of it. Does the scale still balance horizontally? If not, add sufficient tin-foil to one or other pill-box (Fig. 8).

Remove the two boxes and put 20 gm. into each. Hang one box 40 cm. from O on one arm. Find where the other box must be hung that the scale may balance horizontally: it has to be placed 40 cm. from O on the other arm. Now hang the first box in a new position: again, the second box must be placed at an equal distance from O to maintain equilibrium. It is clear that *equal weights will only balance one another when at equal distances from the point of support.*

Again, hang the two boxes 40 cm. from O, and on opposite sides of O. Place 30 gm. in one box, and add gram weights to the other until there is equilibrium. Is it not necessary to add an equal weight, viz. 30 gm.? Repeat with some other weight in the first box. It becomes clear that *two substances must have equal weights if they balance one another when at equal distances from the point of support.*

Suppose we desire to find the weight of some copper turnings. Place them in one box at A. Place the other box at B, such that the length BO = the length AO. Add standard masses or "weights" to the box until there is equilibrium. Then we may conclude that the weight of the copper is measured by the amount of the standard weights in the box.

Of all the instruments which are at the present day in the service of the chemist the balance is the most important. To-day the balance is to the chemist as his very right hand, and without it the rapid progress of the last century would have been impossible. The use of the balance has placed chemistry among the exact sciences, and so far as may be chemical problems are now put to the test in the laboratory, and not referred for an answer to the logician in his study. How different our opinion of the balance is from that of Jean Rey, a shrewd physician, who said in 1630:

"I affirm that the examination of weights which is made by the Balance differs greatly from that which is made by the reason. The latter is only employed by the judicious, the former can be practised by the veriest clown. The latter is always exact; the former is seldom without deception."

The balance is a delicate instrument, and Jean Rey's clown cannot be allowed to play with it, or it will soon be injured.

The Student's Balance.—We shall now describe the balance to be used in almost all the experiments and problems described in this book.

It consists of a beam A B (Fig. 9) which can be raised from or lowered on to a knife edge at C by the handle D. At its two

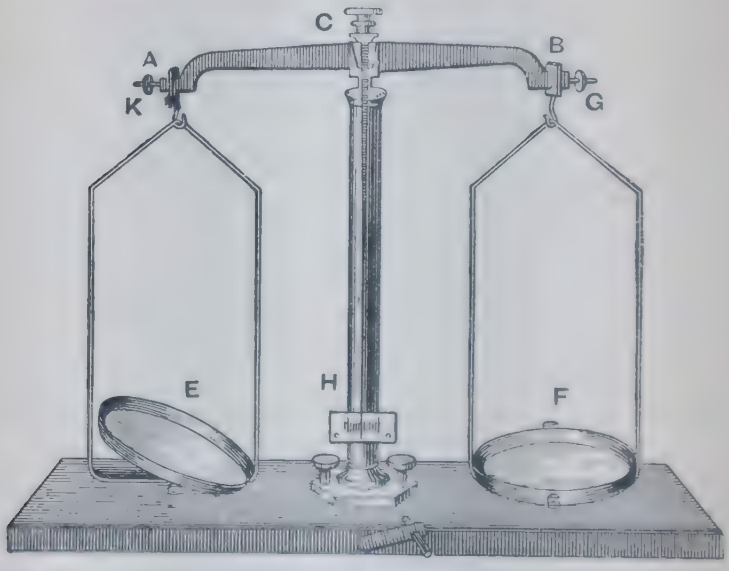


FIG. 9.

ends are suspended pans E and F. A needle H is attached rigidly at right angles to the beam, and its point is in front of a graduated scale. There are screw nuts at K and G by means of which the lengths of the arms can be slightly altered. The nuts K and G should be so adjusted that when the pans are empty the needle swings over the same number of degrees to the right and the left of the zero point at the centre of the scale.[1]

Remove the pans from a balance. Examine the means of

[1] Several firms sell balances of this type. For laboratory work it is recommended to obtain them fitted with glass cases, whereby the balance is protected from the disturbing influence of air currents, and the still more dangerous effects of corrosive chemical fumes.

A Rough Balance.—It is often only desired to know quickly the approximate weights of substances. A very convenient balance is Salter's Letter Balance, No. 11, which will carry 1000 grams. (G. Salter and Co., West Bromwich, 55/6d.)

Another convenient balance is one built on what is known as the French type.

support, including the knife edges. Make a drawing of the way in which the pan is supported.

The Weights.—In a set of weights from 1 centigram up to 50 grams the following individual weights are usually found :—

Brass	50	20	20	10
	5	2	2	1
Flat White Metal	0·5	0·2	0·2	0·1
	0·05	0·02	0·02	0·01

Notice that there are just enough weights to add up to any sum from ·01 to 111·10 grams.

For example, if a substance weighing 67·89 grams were weighed, the weights would be made up as follows :—

$$\left\{\begin{array}{l} +\ 50 + 10 \\ \quad +\ 5 + 2 \\ \quad\quad +\ ·5 + ·2 + ·1 \\ \quad\quad\quad +\ ·05 + ·02 + ·02 \end{array}\right.$$

The gram weight and those greater than the gram are usually made of brass.

The Smaller Weights, which are always used in accurate weighing, are differently marked by different makers. They are made of aluminium, nickel, or platinum.

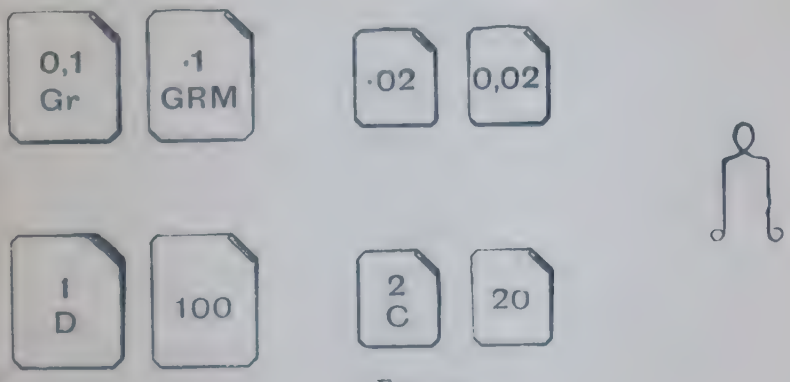


Fig. 10.

The smaller weights shown above in Fig. 10 represent :—

·1 gram	·02 gram	·01 gram
= 1 decigram	= 2 centigrams	= 1 centigram
= 100 milligrams	= 20 milligrams	= 10 milligrams.

* **The Rider.**—The 01 gram or centigram weight is often a small bent wire called a *rider*, because it can ride astride the

balance beam. When placed at different points along the beam, its effective weight changes, and makes it possible to weigh to milligrams or even to tenths of a milligram. The beam of an accurate balance is divided into ten equal spaces. When the rider is put on the outermost of these, it is worth its full weight, ·01 gram. On the middle division it acts as ·005 gram, and on the innermost division as ·001 gram.

EXERCISE.—Remove all the flat weights from the box. Arrange them on the squares drawn in Fig. 11. Ask your

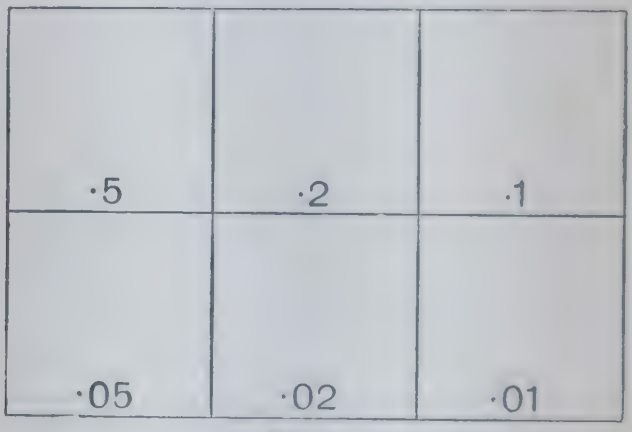


FIG. 11.

teacher whether you are right. Rearrange the weights in their box. They should be in order of magnitude. Never force a weight into a hole too small for it. Give each weight a separate space if you can. The bent-up corner should be at the upper right-hand corner. Ask your teacher whether this has been properly done.

The Method of Weighing.—The following points should be observed :—

1. Always turn the handle D gently, and right over from one side to the other.

2. See that the pointer is swinging equally upon both sides of the scale, when the pans are empty. [Why?]

3. Stop the swinging of the balance whenever the object or the weights are added to or removed from the pans. [Why?]

4. Always place the object to be weighed in the left-hand pan, and the weights in the right-hand pan. [Why?]

5. The position of the observer should be opposite the centre of the scale. [Why?]

6. First add the largest weight necessary, and then try the smaller ones in *descending order of magnitude, omitting none*, but removing those that are too heavy before substituting lighter ones, and arranging the weights on the pan in order of magnitude, with the larger in the centre.

7. When the object is exactly balanced, write down the weights missing from their places in the box, and also the weights on the balance pan, and see that the two agree.

8. Never weigh anything which has been heated, until it is cold.

Exercises

Weighing

1. Find whether the balance swings true, the pointer vibrating to equal distances on each side of the middle point of the scale. Probably it does not. Take a square decimetre of paper. Tear it in two. Letter one half A. Tear the other half in two. Letter one part B. Go on bisecting: letter the weights so made A, B, C, ... H. Use these weights to counterpoise the balance until it does swing true.

2. If the error of the balance is less than a centigram, use a centigram rider, and without any other weight find what weight would adjust the balance.

3. Pick out weights amounting to 87·65, arranging them neatly on the balance pan, the largest in the centre.

4. Add up the weights on the balance pan.

5. Add up the spaces in the box.

6. Take a set of English *grain* weights, 600 grains to 10 grains, and find the weight of each in grams to the nearest decigram *next below the true weight.* Write answers thus—11·9 +.

7. Repeat exercise 6 to the nearest centigram *next below the true weight.*

8. Cut a piece of lead to weigh exactly 10 grams.

9. Cut a piece of lead foil to weigh exactly 1 gram.

10. Weigh newly minted French (or English) silver coins.

11. Weigh the glass stopper[1] provided.

12. Weigh the copper cylinder[1] provided.

[Further exercises in weighing will be found in the succeeding chapters. Excellent exercises will be afforded by determinations of relative density.]

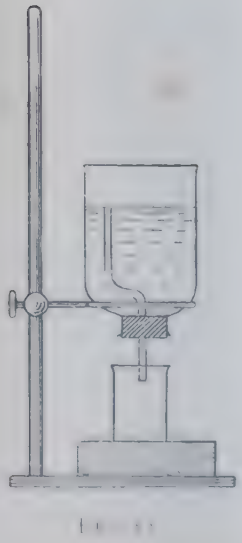
[1] The teacher is recommended to keep a stock of glass stoppers and copper cylinders of ascertained weight.

CHAPTER IV

THE MEASUREMENT OF THE VOLUME OF A LIQUID

The Units of Volume.—These are the **litre** for large volumes, and the **cubic centimetre** for small volumes. The litre is equal to 1000 cubic centimetres.

It is very important to remember that **1 c.c. of water weighs 1 gram.**[1] Consequently, we may regard any given number of grams of water as occupying a volume expressed by the same number of c.c. For example, 1 litre or 1000 c.c. of water will weigh 1000 grams. The convenience of this relation will be learnt from the following experiment.



EXERCISE 1. **To find the Volume of a Pebble.**—We shall know its volume if we can find out how much water it can displace. Take a small bell jar with an open neck. Fit into the narrow end a rubber stopper, through which is passed a narrow glass tube, bent as shown in Fig. 12. Fix it in a clamp. Place a tumbler underneath, and pour water into the jar until it is full up to the mouth of the tube. Tie a thread round the pebble. Weigh a beaker. Place it below the lower end of the tube. Carefully lower the pebble into the water

[1] This is strictly true only at 4° C. The average temperature of a room is about 15° C. However, since water expands slightly between 4° C. and 15° C., 1000 c.c. of water will at 15° C. weigh slightly less than 1000 grams.

until it is quite immersed. The water overflows, and is caught in the beaker. Weigh again. Example—

Weight of beaker...	8·53 gm.
Weight of beaker + water which overflowed...	29·97 ,,
∴ Weight of water which overflowed	21·44 ,,
and Volume ,, ,, ,, ...	21·44 c.c.
Hence, volume of pebble	**21·44 c.c.**

It is now evident that we can readily find the volume of a flask by first weighing it when empty, and secondly when filled with water. For the increase in weight expressed in grams will measure its volume expressed in cubic centimetres.

EXERCISE 2. **To find the Volume of a Flask** up to a mark on the neck. Gum a label on the neck, and make a horizontal mark on the label. Weigh the flask. Fill it with water up to the neck. If the top surface of the water in the neck of the flask be observed, it is seen to be curved (Fig. 13). This curved surface is called the **meniscus**. If there is a horizontal mark on the neck of a flask, it is said to be full up to the mark when the *lowest* point of the meniscus is on the level of the mark. Carefully bring the meniscus to the mark. Enter results thus :—

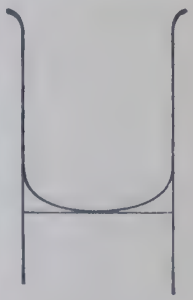


FIG. 13.

Weight of flask	24·21 gm.
Weight of flask filled with water...	106·42 ,,
Hence, weight of water in flask	82·21 ,,
∴ Volume of flask	**82·21 c.c.**

Measuring Flasks.—In carrying out experimental work it is found convenient to have flasks with which to measure out a definite volume of liquid, for instance, 100, 250, 500, 1000 c.c.

EXERCISE 3. **To mark a Flask to contain 100 c.c.**—Take a dry 3-ounce flask with a narrow neck. Weigh empty. Then add water till there is an increase in weight of 100 grams. Allow the flask to stand until any water on the side of the neck has drained down. With a sharp file moistened with water

scratch a horizontal line on the neck to mark the lowest point of the curved surface bounding the water and the air.

The measuring flasks most often used are litre, $\frac{1}{2}$-litre, $\frac{1}{4}$-litre, and 100 c.c. flasks (Fig. 14). They are usually fitted with ground-glass stoppers, and when filled up to the circular line etched upon their necks contain at 15° C. the volumes which are marked upon them.

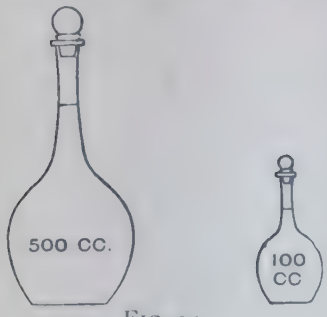


FIG. 14.

Measuring Jars.—These are tall cylindrical vessels, graduated in cubic centimetres, and holding 100, 200, 250, 500, or 1000 c.c. (Fig. 15).

The student will understand how to read the level of a liquid in a measuring jar after studying the use of a burette.

Pipettes.—Flasks and measuring jars measure the liquids they *contain*, but it is extremely useful to have vessels which will measure out, or *deliver*, accurately a small definite volume of liquid.

FIG. 15.

A pipette is a glass tube open at both ends, upon which a cylindrical bulb has been blown (Fig. 16). There is a circular mark on the stem. When filled with liquid up to the mark, the pipette will deliver exactly that volume which is marked upon it.

EXERCISE 4. **To use a Pipette.**—Place the pointed end of the pipette in water. Suck up the water till above the mark. Quickly close the upper end with the moist thumb. Pressing gently, rotate the pipette so that the water very slowly flows out until the meniscus reaches the mark. Then press tightly. The flow ceases. Transfer the pipette to a beaker, remove the thumb, and let the liquid flow out. As soon as the pipette appears to be empty, let it drain for five seconds, with the tip just below the surface of the water in the beaker. The drop of water which is still in the pipette is left, and not blown out.

Exercises

Use of Pipette

1. Suck up water in a pipette. Set to the mark.

2. Take a 10 c.c. pipette and a small beaker (2 oz.). Find how many 10 c.c. of water can be added to the beaker.

3. Take a 10 c.c. pipette and a 100 c.c. flask. Find the level of the water after you have added 10 c.c. of water to the flask 10 times.

4. Repeat No. 3, using a 100 c.c. measuring jar in place of the flask.

5. Weigh an empty beaker. Measure into it 5 quantities of water of 10 c.c. each by means of a pipette. Weigh again. What is the gain in weight?

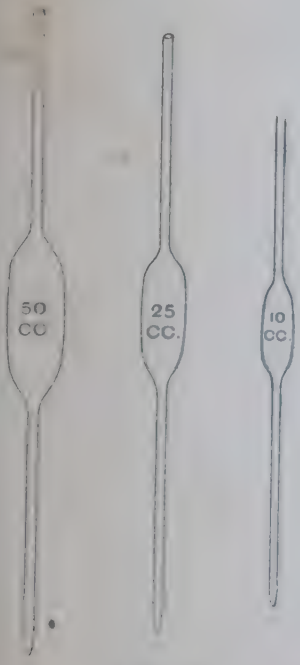


Fig. 16.

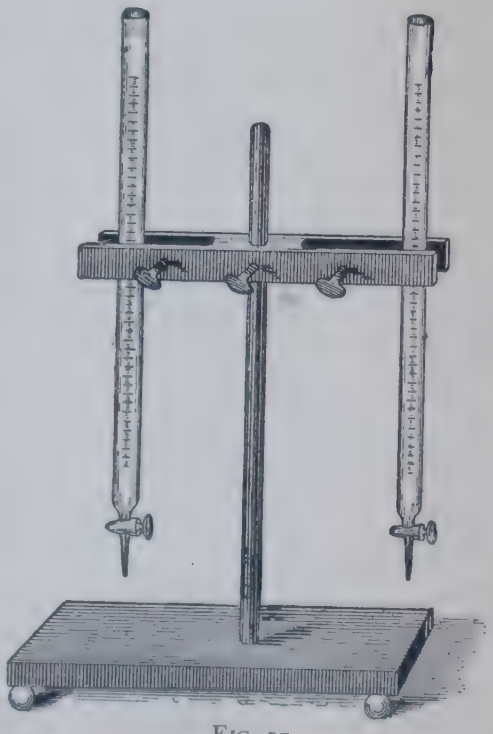


Fig. 17.

Burettes. A burette is a glass tube with a glass tap at the lower end, used to deliver any desired volume of liquid (Fig. 17). The tube is graduated in cubic centimetres and tenths of a cubic centimetre. In place of a glass tap a burette may be fitted with a glass jet attached by an india-rubber tube, on which is

placed a spring clip ; such a burette is more suitable for school use.

EXERCISE 5. **To prepare a Burette for Use.**—Fix the burette upright in a clamp. Place a spring clip upon the rubber tube. Pour water into the tube through a small funnel, until its level is above the top of the scale. Remove the funnel. Open the clip, and let some water run out, so that no air is left in the burette below the clip. The burette is now ready to deliver a measured volume of water.

Examine the divisions and figures upon the burette. Draw the portion from 41 to 43 four times the actual size.

Burette-Reading.—The surface of a liquid in a burette, like that in a flask, is curved, and is called the *meniscus ;* it is always the *lowest point* of the meniscus which is measured. Note the following points :—

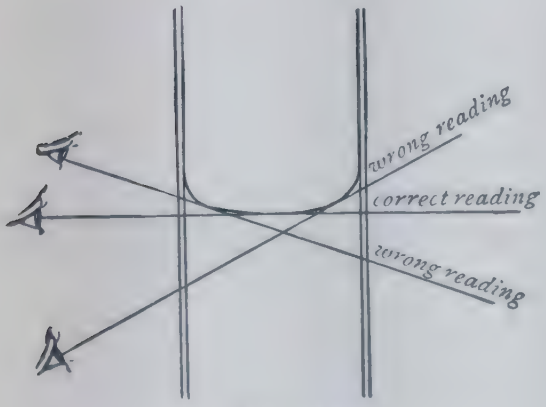


FIG. 17 A.

(1) *The eye must be on the same level as the curve*, otherwise the curve is seen displaced above or below the true position (see Fig. 17 A). The level position may be obtained thus. Stand a few paces back from the burette, and decide upon some object behind the burette, which appears to be on the same level as the meniscus. Then go up to the burette, and place the eye so that the meniscus is seen exactly on a level with the object selected.

(2) The extreme edge, or *lowest position*, of the meniscus is, however, sometimes hard to see, but it can be made clear. On looking into an aquarium from below the surface-line gold-fish are seen reflected above the surface. This phenomenon of reflection from the inside of a surface is called *internal reflection*, and it is this which may be made use of to render the meniscus distinct. If the background is not bright, the meniscus is made clear by holding a white object *just behind and just below the surface*. The edge of a card held behind the burette

tube answers admirably. At night, however, when the background is dark, a gas flame may be put on the bench below to illuminate the meniscus.

(3) The graduations are decimal, each fifth and tenth line being longer than the rest. Notice the figures opposite the divisions. Find whether they count upwards or downwards. Find how many c.c. or what decimal of 1 c.c. each scale-division stands for.

Examples VI

Burette-Reading

1. Read the position of the liquid in the figure of a tube shown on next page (Fig. 18), taking the two long lines W and X to represent respectively :—

	A	B	C	D	E	F	G	H
	30	40	7	800	120	130	250	250
and	20	50	6	900	100	150	200	300

For instance, the two lines W and X represent 30 c.c. and 20 c.c. respectively in question A, 40 c.c. and 50 c.c. in question B, and so on.

2. Sometimes the lowest point of the meniscus is not exactly opposite to any of the scale-divisions; tenths or other fractions of a division should then be *estimated* (see page 9).

Read the position of the liquid shown in Fig. 19, estimating the position of the curve in tenths of a scale-division, taking the two long lines Y and Z to represent respectively :—

	J	K	L	M	O	P	Q	R
	1	14	10	30	40	700	800	0
and	2	13	11	20	50	600	900	1000

Exercises

Burettes, Measuring Jars, and Pipettes

(*a*) *Burettes*

(1) Pour water into a burette. Set it exactly at 10 c.c.

(2) Pour some water out of the burette. Read the new position of the liquid.

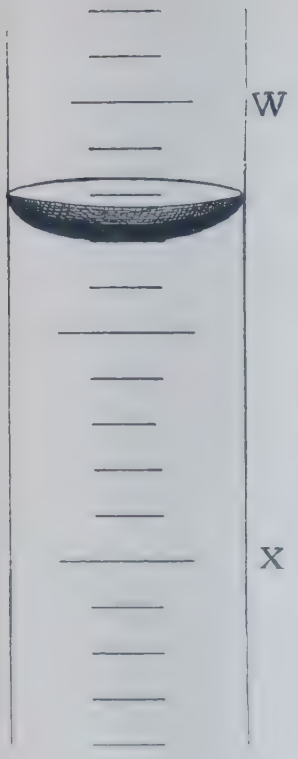


FIG. 18.

FIG. 19.

(3) Choose a 100 *grain* pipette and a 50 c.c. burette, graduated to tenths of 1 c.c. Set the liquid in a burette to 50. Measure 100 grains of water in the pipette, and run it into the burette. Read the level now. Run in another 100 grains, read again, and so on until you read the top of the burette. Take the difference between each two consecutive readings. These differences ought to be equal, for each should be the number of c.c. occupied by 100 grains of water. If the differences are not nearly equal, find out where you have made a mistake. Repeat the exercise with still greater care, estimating tenths of a division.

State your results thus :—

EXPT. 1. Measuring to ·1 c.c.

		Differences.
0.	50	
		. . 6·3
1.	43·7	
		. . 6·5
2.	37·2	
		. . 6·4
3.	30·8	
		. . 6·5
4.	24·3	
		. . 6·5
5.	17·8	
		. . 6·4
6.	11·4	
7.	. . . . .	4·69

EXPT. 2. Estimating ·01 c.c.

	Differences.
50	
	. . 6·49
43·51	
	. . 6·46
37·05	
	. . 6·49
30·56	
	. . 6·46
24·10	
	. . 6·48
17·62	
	. . 6·48
11·14	
	. . 6·45
4·69	

4. Pour water into a burette. Set it exactly at 0 c.c. Weigh an empty beaker. Run into it exactly 5 c.c. Weigh. Run in another 5 c.c., weigh again, and so on until you have run out 50 c.c. Take the difference between each two consecutive weighings. These differences should be equal if the burette is correctly graduated.

5. You are provided with a small bottle. Find its volume by means of a burette.

(*b*) *Measuring Jars*

6. Make a drawing, twice the real size, of the portion of a measuring jar from 60 to 80. Show the graduations and the figures.

7. Add water to a measuring jar until the level of its surface is exactly at 74 c.c.

8. Read the level of the water in the measuring jar provided.

9. You are provided with a measuring jar and some marbles. Devise a method of finding their volume. Find also the average volume of 1 marble.

(*c*) *Pipettes* (see also p. 33)

10. Find the volume of a 100 grain pipette as follows: Weigh an empty beaker. Fill the pipette with water, and run the water into the

beaker. Repeat three times. Again weigh. Calculate the weight in grams of 100 grains of water.

11. Recalibrate the pipette provided, thus:—Gum a strip of paper 1½ inch long upon the stem of a pipette. Find the volume up to the bottom, and also to the top edge of the paper, proceeding as in No. 10. Calculate where the mark should be placed upon the stem, and mark it on the paper.

CHAPTER V

THE MEASUREMENT OF TEMPERATURE

Heat and Temperature.—Every one can attach some meaning to the word *Heat*. The sensation of heat is familiar to all. And yet we shall often make mistakes in judging whether a substance is hot or cold by our sensations. For the same room may at the same time appear hot to one man and cold to another. Indeed, the two hands may judge differently of the same substance; test this by placing one hand for a few seconds in hot water and the other in cold, and then place both together in lukewarm water: one hand feels cold and the other warm. It is clear, then, that the sensations of heat and cold in our bodies do not necessarily tell us anything of the actual state of a substance with regard to heat or cold.

Nevertheless there is no room for doubt that different states of a body with regard to heat and cold can exist, and we shall proceed to consider how these different states or conditions can be measured. This we can do without stopping to find out exactly what heat itself is.

Temperature.—If a poker be placed in a fire, the end outside the fire quickly becomes warm, and we understand that heat has passed from the hot coals into the poker. If the red-hot poker be placed in a bucket of cold water, the poker will be rapidly cooled, and we say that heat has passed from the poker into the water. Many instances are met with every day of heat passing from one body to another; such bodies are said to be at different *temperatures*, and the heat passes from the body which is said to have the higher temperature to that which is said to have the lower temperature.

A difference in temperature between two bodies may be compared to a difference in level between the water in two cisterns. When two bodies are so placed that heat can flow between them, the flow of heat will always take place from the body which has the higher temperature to that which has the lower temperature, just as water will always flow from a higher level to a lower. Indeed, temperature may be defined as *that condition of a body upon which depends the flow of heat to or from other bodies.* It is important also to observe that the passage of heat from a hot to a cold body will continue until they reach a common temperature, just as water is said to find its own level.

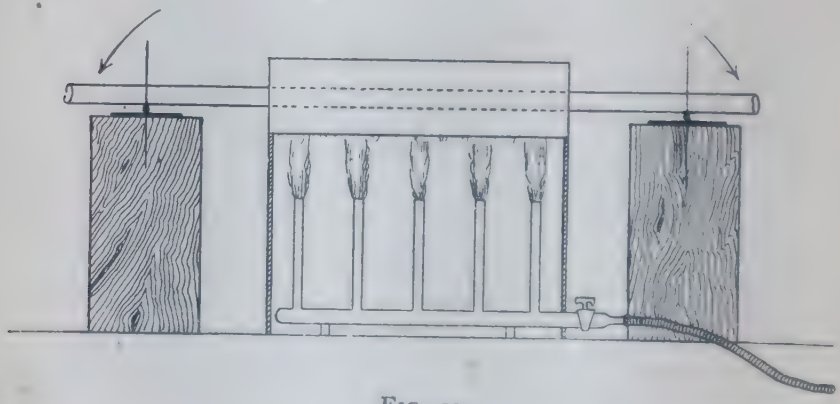


FIG. 20.

Preliminary Idea of a Thermometer.—When heat is applied to any body it may produce many effects, but none is more noticeable than change in volume. In by far the greater number of cases a body expands or increases in volume when it is heated.

EXPT. 1. **To show that Iron expands when heated.**[1]—Support an iron gas pipe at its two ends upon two needles, resting upon horizontal glass plates, and let the central portion of the pipe lie within (but not touching) a gas furnace (Fig. 20). Fix to each needle a light straw to serve as an index. Now light the furnace. In a short time each needle begins to roll outwards turning the straw. This movement must be caused by the lengthening (expansion) of the bar. Note that the hotter the pipe becomes the greater is the expansion, and the consequent

[1] Lecture Table Experiment.

movement of the straws. Allow it to cool. Note how the needles roll back.

EXPT. 2. **To show that Water expands when heated.** —Take a glass flask. Select a rubber stopper, with one hole, which fits it closely. Pass a narrow glass tube, 50 cm. in length, through the cork, so that the end of the tube is flush with the lower surface of the stopper (Fig. 21). Fill the flask quite full with water coloured with indigo. Push the stopper with the tube into the neck of the flask so that all air is expelled, and the liquid rises a short way up the tube. Mark the level of the liquid with a piece of gummed label. Immerse the whole of the flask in a vessel containing hot water. What do you observe? The liquid in the tube first falls slightly; then it rises steadily. Why? The fall at first may have been due to a contraction of the water, or to an expansion of the glass flask. The heat had hardly time to reach the water, so probably the fall was due to the flask expanding. The second change—the slow steady rise—is due to the gradual heating and expansion of the water within the flask.

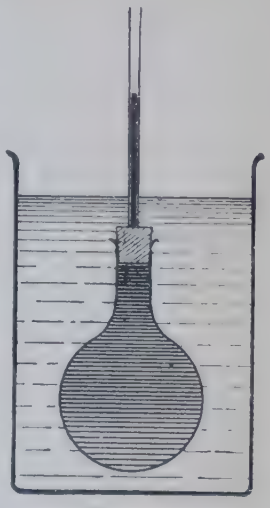


FIG. 21.

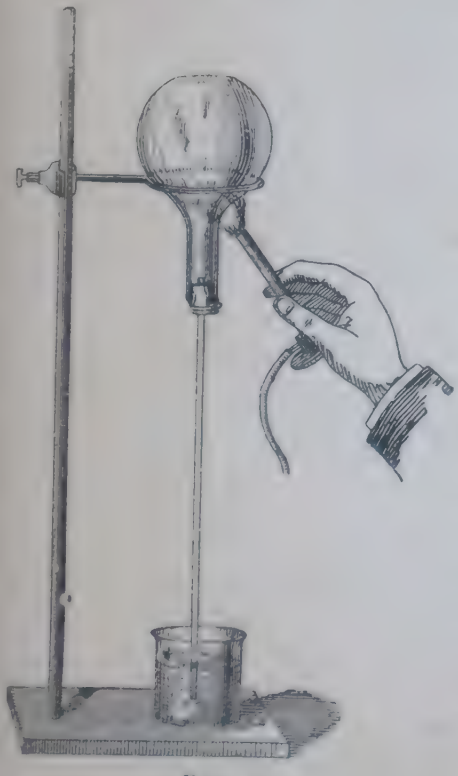


FIG. 22.

EXPT. 3. **To show that Air expands when heated.** —Close a glass flask with a rubber stopper through which is passed a long glass tube. Invert the flask and support it so that the open end of the tube dips under coloured water contained in a beaker (Fig. 22).

Warm the flask first with the hand. A few bubbles escape. Then warm the flask with a flame. The enclosed air bubbles out much faster. Remove the source of heat. Water rises up the tube, showing that the air contracts on cooling.

In these three experiments we have illustrations of the expansion by heat of a solid, of a liquid, and of a gas. In each case the expansion indicates the flow of heat into them from sources which are at higher temperatures. Hence, if we find that a substance is expanding in volume, we may conclude that heat is probably passing into it from something else which must be at a higher temperature.

Fig. 23.

The simple instrument described in Experiment 2 may be used to compare the temperatures on different days, or in different parts of a house, provided in each case the instrument is left sufficiently long for the passage of heat to continue until the instrument and the room have acquired a common temperature. If the glass tube be graduated—that is to say, divided into equal divisions and numbered—temperatures may be stated in terms of this scale. Such an apparatus is a simple instrument for recording and measuring temperatures—in other words, it is a rough **thermometer.** It may be remarked that Galileo in the year 1597 was probably the first to employ a thermometer.

The Mercurial Thermometer

Exercise.—1. Examine a mercurial thermometer. Draw the bulb and a portion of the stem four times the actual size.

2. Try the effect of holding the bulb in the hand, and of placing it in the mouth.

The ordinary mercurial thermometer is really just such an instrument as was used in Experiment 2. Mercury, however, is used as the liquid instead of water. The principal advantages of mercury over water are, that it does not *freeze* until a very low temperature, and secondly, that it does not *boil* until a very high temperature. The mercurial thermometer consists of a very fine thick-walled glass tube, which ends in a round or cylindrical bulb (Fig. 23).

The Fixed Points on a Thermometer.—In order that a thermometer may be of practical use, its stem must be

graduated. And, further, all thermometers should be marked so as to show the points at which the mercury would stand were all exposed alike to one or two particular temperatures. Otherwise, it would be impossible to compare the readings of one thermometer with those of another.

In graduating the thermometer it is customary in the first place to mark upon its stem the points at which the mercury stands when it is heated to two standard temperatures. It was Newton who suggested that for these **fixed points**, as they

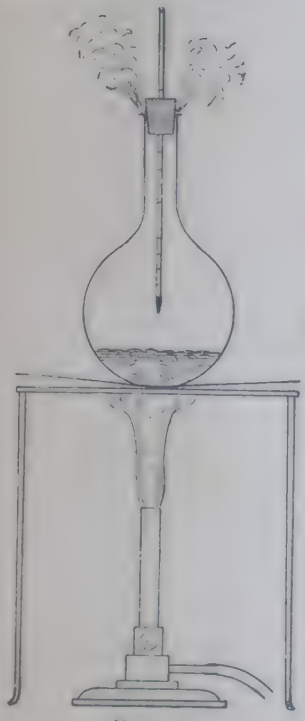


FIG. 24.

FIG. 25.

are called, the melting temperature of ice, which is found to be very constant, and the temperature of steam arising from boiling water, which also, as will be shown later, is constant so long a the pressure of the air remains the same, should be employed.

EXPT. 4. **To mark the Boiling Point on a Thermometer.**—Take a large flask containing some water, and fit it with a cork. File two grooves down the side of the cork, and pierce it with a hole through which a thermometer can be thrust, so that its bulb is just above the surface of the water (Fig. 24).

Boil the water vigorously. The thermometer is now surrounded by steam. Observe carefully the point to which the mercury rises. This point is marked one hundred degrees or 100° on a Centigrade and 212° on a Fahrenheit thermometer.

EXPT. 5. **To mark the Freezing Point on a Thermometer.**—Fill a funnel with pounded ice. Support it on a tripod. Plunge the thermometer into the mass (Fig. 25). Observe the point at which the mercury becomes stationary. This is marked 0° on a Centigrade and 32° on a Fahrenheit thermometer. Is your thermometer marked correctly?

Scales of Temperature.—The difference in temperature between these two fixed points is large. It is customary to subdivide the interval between them according to one or other of three scales, named the Fahrenheit, Centigrade, and Réaumur scales; we need only consider the first two.

On the Fahrenheit scale the freezing point is marked 32° and the boiling point 212°, while on the Centigrade scale they are marked 0° and 100° respectively (Fig. 26). The interval between these points is therefore divided into 180 equal divisions or degrees on the Fahrenheit and into 100° upon the Centigrade scale.

Hence,	180	Fahrenheit degrees	=	100	Centigrade degrees.	
or	9	,,	=	5	,,	
Therefore,	1	,,	=	$\frac{5}{9}$	,,	,,
and	$\frac{9}{5}$	,,	=	1	,,	,,

The position marked 0° on a scale is often called the **zero** of the scale. On the Fahrenheit scale zero is 32° below freezing point.

Degrees below 0° are marked by the *negative* sign: thus −10° C. means 10° below 0° C.

To Read the Thermometer.—Although the interval between the fixed points on a Centigrade thermometer is divided into only 100 degrees, it is nevertheless possible by means of the eye to mentally subdivide each degree into tenths, and therefore to read any given temperature to the nearest *tenth* of a degree.

Fig. 27 shows a piece of a thermometer stem. The graduations stand for single degrees. The end of the mercurial column lies between 11° and 12°. Subdividing the eleventh degree by the eye, the reading is 11°4.

Also observe the following points :—

1. The eye should be at right angles to the scale at the point reached by the mercury (compare p. 34).

2. The thermometer should be read while its bulb is *in* the substance of which the temperature is being taken.

3. The figures count upwards.

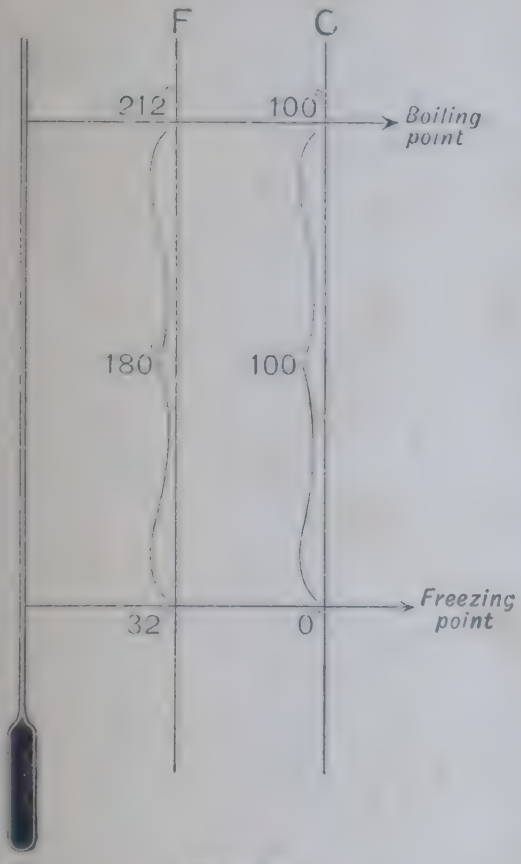


FIG. 26.

FIG. 27.

EXERCISES

1. Fill a tumbler with cold water from the tap. Place a C. thermometer in the water. Read the temperature. Change the water and read again.

2. Take some cold water in a beaker, and place a C. thermometer in it. Heat the water over a small flame. When the temperature reaches 20° C., dip your fingers into the water, and describe your impression. Continue to heat the water, and when the temperature is 40° C. and 60° C. try it in a similar way.

3. Read the thermometers placed about the room and up the wall.

4. What is the lowest temperature that can be read with the thermometer provided? What is the highest temperature?

5. Read the position of the arrow on a clinical thermometer.

Conversion of Thermometric Scales.—

(1) Find the Centigrade reading corresponding to 80° F.

80° F. is (80° − 32°) or 48° F. above the freezing point.

48° F. is $48° \times \frac{5}{9}$ C. or 26°·6 C. above the freezing point, which is 0° on the C. scale.

Hence, the required reading is 26°·6 C.

(2) Find the Centigrade reading corresponding to 20° F.

20° F. is (32° − 20°) or 12° F. below the freezing point.

12° F. is $12° \times \frac{5}{9}$ C. or 6°·6 C. below the freezing point, which is 0° on the C. scale.

Hence, the required reading is −6°·6 C.

(3) Find the Fahrenheit reading corresponding to 15° C.

15° C. is $15° \times \frac{9}{5}$ F. or 27° F. above the freezing point, which is 32° on the Fahrenheit scale.

Hence the required reading is (27° + 32°) F. or 59° F.

***Mechanical Interpolation.**—The conversion of thermometer scales is only one instance of the mathematical process called *interpolation*. Other instances occur in making out railway time-tables, in arranging prices and sizes of goods, and in the use and preparation of logarithmic, astronomical, and other tables (see Chapter XVIII.).

EXAMPLES VII

Interpolation—Conversion of Thermometric Scales

(i) * INTERPOLATION

1. In a certain street the lamp-posts and the houses do not count from the same point. The third lamp-post is opposite No. 21, and the seventh lamp-post is opposite No. 45. What number is opposite the tenth lamp-post?

2. On a certain railway journey by express K, L, and M are passed at 12.3, 12.21, and 12.27. A slower train passes K and M at 1.4 and 1.40. At what time does it pass L?

3. If the sun sets at 6.10 P.M. on September 16, and at 5.54 on September 23, find the time of sunset on September 21.

4. A solution containing 40 per cent. sulphuric acid has the density 1·306, and a 50 per cent. solution has the density 1·398. What is the density of a 46 per cent. solution?

(ii) Conversion of Thermometric Scales

What Fahrenheit temperatures are equal to

5. 25° C. 6. 5° C. 7. 80° C.?

What Centigrade temperatures are equal to

8. 50° F. 9. 95° F. 10. 122° F.?

Express on the Fahrenheit scale—

11. 20° C. 12. −5° C. 13. 4°·5 C.

Express on the Centigrade scale—

14. 40° F. 15. 14° F. 16. −4° F.

17. On the Réaumur scale the freezing point is 0° and the boiling point 80°. What Réaumur temperature is equal to 59° F.?

18. What Fahrenheit temperature is equal to 16° R.?

Exercises

Thermometry

1. Fill a beaker with cold water from the tap. Place a F. and a C. thermometer in the water. Read the lowest temperatures they show, as nearly as possible at the same time. Change the water, and again read. Repeat four more times. Enter thus :—

F.	68°·9						
C.	15°·0						

Convert the last C. reading to the F. scale. Does the number agree with the last F. reading?

2. Warm the water in the beaker, and when the C. thermometer reads about 40°, remove the flame, stir, and read both thermometers accurately, and as nearly as possible at the same time. Convert the C. reading to the F. scale. Does the result agree with the F. reading?

3. Heat the water again, until the C. thermometer reads about 75°, and then read both thermometers. Convert the F. reading to the C. scale. Does the result agree with the C. reading?

4. Now boil the water, and place the C. thermometer in it. Arrange two columns in your note-book headed **Time** and **Temperature**, thus—

Time.			Temperature.
3 hr.	20 m.	0 sec.	100°·2 C.
3	20	30	
3	21	0	
3	21	30	

Look at the second-hand of your watch, and when it is at the minute read the thermometer and at once extinguish the flame, but leave the thermometer in the water. Put down the temperature of the boiling water opposite to the first time reading. Then, at intervals of half-minutes, take the readings of the thermometer to the first decimal place. Take 16 readings. Plot the results on squared paper. (See Chapter XVIII.).

5. Half fill a flask with water. Add about 10 grams of common salt. Heat the flask until the water boils. Hold the bulb of a C. thermometer in the steam *above* the boiling liquid. What is the highest temperature it records?

Then place the bulb of the thermometer *in* the boiling liquid. Is the reading the same?

6. Pound some ice in a mortar. Add a small handful of common salt. Mix thoroughly. Read the lowest temperature reached.

CHAPTER VI

THE MEASUREMENT OF THE RELATIVE DENSITIES OF SOLIDS AND LIQUIDS

I. THE RELATIVE DENSITIES OF LIQUIDS

No one can fail to be aware that there are many different kinds of liquids. At the mention of mercury, water, oil, or methylated spirits, we may also perhaps admit that they are not all equally heavy. But how can we actually compare the weight of one liquid with that of another? Obviously by weighing in turn one and the same volume of each liquid. We know that 100 c.c. for instance of water weigh 100 grams. Suppose we find the weight of 100 c.c. of other liquids.

We can readily measure out equal volumes of different liquids by means of a pipette.

EXPT. 1. **To find the Weight of 100 c.c. of Methylated Spirits.**—Take a small dry beaker. Weigh. Measure into it by means of a pipette 20 c.c. of methylated spirits. Weigh. State thus:—

Weight of beaker	= 15·41 gm.
Weight of beaker + 20 c.c. spirits	= 31·45 ,,
∴ Weight of 20 c.c. spirits	= 16·04 ,,
∴ Weight of 100 cc. spirits... ...	= **80·20** ,,

EXERCISES.

Find the weight of 100 c.c. of—

(1) milk.
(2) oil of turpentine.
(3) chloroform.
(4) hydrochloric acid.
(5) sulphuric acid.
(6) sea water.

Compare your results with the following numbers, which were obtained in the same way :—

Methylated spirits ...	80·2 gm.	Milk	102·9 gm.
Oil of turpentine ...	86·3 ,,	Hydrochloric acid ...	114·1 ,,
Benzene	88·2 ,,	Chloroform	147·1 ,,
Sea-water	102·5 ,,	Sulphuric acid ...	178·2 ,,

It is evident that equal volumes of these liquids contain different amounts of matter. This is usually expressed by saying that some liquids are **denser** than others, or that they possess different **densities.**[1]

Now, it is convenient to compare the densities of different liquids with that of some one liquid, and water has been chosen as the standard liquid. We can at once obtain from the results of our experiments numbers which will express the **density** of any liquid **relative** to that of water.

Definition of Relative Density.—

$$\text{Relative density or Specific gravity} = \frac{\text{Weight of a substance}}{\text{Weight of an equal volume of water}}$$

For instance :—

$$\text{Relative density of milk} = \frac{\text{weight of 100 c.c. milk}}{\text{weight of 100 c.c. water}} = \frac{102{\cdot}9}{100} = \mathbf{1{\cdot}029.}$$

In this way we find the following relative densities, or specific gravities, as they are often termed :—

Methylated spirits	0·802	Milk	1·020
Oil of turpentine	0·863	Hydrochloric acid	1·141
Benzene	0·882	Chloroform	1·471
Water	1·000	Sulphuric acid	1·782
Sea-water	1·025		

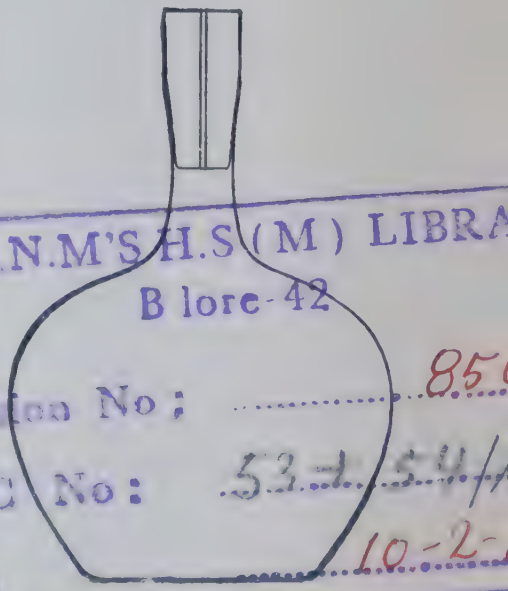
[1] Definition of Density.—*The density of a substance is defined as the mass of unit volume.*

Hence, for example, from the results above—

Density of methylated spirits = 0·802.
Density of sulphuric acid = 1·782.

Observe the meaning that these numbers bear. The statement that the relative density of milk, for example, is 1·029 means that volume for volume, milk is 1·029 times as heavy as water, and that 1 c.c. of milk therefore weighs 1·029 grams.

It is important to observe that determinations with different samples of any one liquid—for example, oil of turpentine—give very nearly the same numbers. This fact leads us to suspect that any definite liquid has a relative density which is a constant property of it (at a given temperature). So long as the liquid is not mixed with any other liquid, its relative density appears to remain constant, and to have a fixed value. Consequently, it is possible to distinguish between liquids which may otherwise resemble one another by a reference to their relative densities. For instance, the adulteration of milk with water, or the addition of water to alcohol, may be detected by finding the change of relative density which results from the admixture.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The Use of a Relative Density Bottle.—Determinations by the above method are rapidly made, and give excellent practice in the use of a pipette. But equal volumes of liquids can be more accurately measured in a Relative Density Bottle (Fig. 28). The glass stopper has a fine hole bored through it, and is ground to fit the neck accurately. The bottle is filled with liquid, and the stopper then lowered gently into place, so that all air bubbles and extra liquid escape through the hole. The top of the stopper and outside of the bottle are wiped dry before weighing. So accurate is this instrument that, if held in the hand for a few seconds, the liquid, expanded by the heat of the hand, will begin to exude through the hole.

These bottles are usually made to contain exactly 25 grams of water at 60° F. But this should be tested by experiment.

Members of a class are recommended to determine by means of this bottle the relative density of some one liquid, *e.g.* methylated spirits or a solution of salt. Results should only differ in the third place of decimals.

CALCULATIONS.—1. Find the weight of 50 c.c. of mercury: relative density of mercury = 13·5

State thus :—

1 c.c. of mercury weighs 13·5 gm.
∴ 50 c.c. ,, ,, 13·5 × 50 gm
or 675 gm.

2. Find the volume of 100 grams of sulphuric acid: relative density of sulphuric acid = 1·85.

1·85 gm. sulphuric acid occupy 1 c.c.
∴ 1 ,, ,, ,, ,, $\frac{1}{1\cdot85}$ c.c.
∴ 100 ,, ,, ,, ,, $\frac{100}{1\cdot85}$ c.c.
or 54·0 c.c.

Relative Densities of some Common Liquids.—

Ether	·73
Alcohol	·80
Benzene	·88
Turpentine	·89
Olive oil	·91
Water	1·00
Milk	1·03
Sea-water	1·026
Glycerine	1·26
Chloroform	1·47
Sulphuric acid	1·85
Mercury	13·59

EXAMPLES VIII

Relative Densities of Liquids

(*a*)

1. Find the weight of 100 c.c. of chloroform: relative density of chloroform = 1·47.

2. Find the weight of a litre of alcohol: relative density of alcohol = 0·80.

3. 50 c.c. of turpentine weigh 43·1 grams. Find the relative density of turpentine.

4. 60 c.c. of a sample of milk weigh 61·4 grams. Find the relative density of the milk.

5. One litre of a certain liquid weighs 1 kilogram. What is the liquid probably?

6. Find the volume of 1 gram of mercury: relative density of mercury = 13·5.

7. Find the volume of 1000 grams of olive oil: relative density of olive oil = 0·9.

8. Find the volume of 1 gram of alcohol: relative density of alcohol = 0·80.

9. Find the volume occupied by 51·3 grams of sea-water: relative density of sea-water = 1·026.

10. Find the volume occupied by 31·5 grams of glycerine: relative density of glycerine = 1·26.

(*b*)

[*Answers to 4 significant figures.*]

11. Two beakers are placed one on each pan of a balance, and shot are added until the beam swings evenly. Then 5 c.c. of mercury are placed in one beaker, and it is found that 67·5 c.c. of water have to be added to the other beaker to restore the balance. Calculate the density of mercury relative to water, and the density of water relative to mercury.

12. A pipette contains 30·86 grams of spirits of relative density 0·81. What is the volume of the pipette?

13. What weight of sulphuric acid of relative density 1·86 will the same pipette contain?

14. 20 c.c. of a liquid A of relative density 1 are added to 20 c.c. of a liquid B of relative density 0·81. What is the weight of the mixture?

15. The mixture referred to in Ex. 14 measures only 38·5 c.c., owing

to contraction having taken place. What is the relative density of the mixture?

16. A flask weighs 25 grams when empty, 125 grams when full of distilled water, and 127·6 when full of sea-water. Find the relative density of the sea-water.

17. How many c.c. of sulphuric acid (1·85) must be measured out to obtain 111 grams of it?

18. Alcohol (·8) and sulphuric acid (1·85) are to be mixed in the proportion 1 : 3 by weight. What volume of alcohol should be taken with 200 c.c. of acid?

19. A salt solution has the relative density 1·025; how many c.c. of water must be added to 1 litre of it to reduce its relative density to 1·020?

20. Find the relative density of a mixture of 100 c.c. water with 100 c.c. spirits of relative density 0·81, assuming that the liquids contract by 2 per cent. of their original volumes on mixing.

21. A 200 c.c. flask weighs 300 grams when full of water. What will it weigh when full of mercury (13·5)?

22. A bottle weighs 20 grams. Full of water at 4° C. it weighs 70 grams. When filled with water at 65° C. it only weighs 69 grams. Find the relative density of the hot water.

23. A tube 30 cm. long weighs 15 grams when empty, and 96 grams full of mercury (13·5). Find in sq. mm. the area of the cross section of the tube.

II. THE RELATIVE DENSITIES OF SOLIDS

All that has been said of liquids, as regards the differences in density which are found to exist, is of course true also of solids. Although it would appear at first sight more difficult to measure the relative densities of solids, there are several ways in which it can be carried out. Remember that it is required to compare the weight of a solid with the weight of an equal volume of water. How can we find the weight of an equal volume of water?

(A) FIRST METHOD: REGULAR SOLIDS

We can easily determine the volume of most *regular* solids, *e.g.* cubes, cylinders, or spheres, by direct measurement and a simple calculation.

EXPT. 2. **To find the Relative Density of a Rectangular Oak Block.**—The volume is readily calculated from the length, breadth, and height of the block. Measure each of

these accurately in cm. and decimal. Also weigh the block. Enter measurements thus :—

length = 5·0 cm.
breadth = 5·0 cm.
height = 2·5 cm.

∴ volume of block = 5 × 5 × 2·5 c.c.
= 62·5 c.c.

∴ weight of an equal volume of water = 62·5 gm.

Also, weight of block = 48·16 gm.

Hence, relative density of the oak $= \frac{48 \cdot 16}{62 \cdot 5}$

$= \mathbf{0 \cdot 77}$.

EXPT. 3. **To find the Relative Density of a Copper Cylinder.**—Volume of a cylinder = $(\text{radius})^2 \times \frac{22}{7} \times$ height. Measure the diameter and height of the cylinder in cm. and decimal. Then :—

height = 3·78 cm.
diameter = 1·56 cm.
∴ radius = 0·78 cm.

∴ area of circular end $= (0 \cdot 78)^2 \times \frac{22}{7}$ sq. cm.

∴ volume of cylinder $= (0 \cdot 78)^2 \times \frac{22}{7} \times 3 \cdot 78$ c.c.
= 7·26 c.c.

∴ weight of an equal volume of water = 7·26 gm.

Also, weight of cylinder = 64·54 gm.

Hence, relative density of copper cylinder $= \frac{64 \cdot 54}{7 \cdot 26}$

$= \mathbf{8 \cdot 88}$.

EXERCISES

Find the relative density of—

(1) cubes or rectangular blocks of different hard woods
(2) an iron or brass cylinder.
(3) a box-wood ball or a large glass marble.

[Volume of sphere $= \frac{4}{3} \times \frac{22}{7} \times (\text{radius})^3$.]

How can we find the volume of a number of glass stoppers, of leaden shot, a heap of sand, or a bundle of steel screws? We shall describe three methods which may be followed in one case or another.

(B) SECOND METHOD: USE OF A MEASURING JAR

EXPT. 4. **To find the Relative Density of some Glass Stoppers.**—Weigh a few glass stoppers. How can their volume be found? They are too irregular for it to be found by mensuration. But we can find how much water the stoppers can displace in a measuring jar. Choose a measuring jar, the narrowest into which the stoppers will easily slide. Partly fill the jar with water, and read the level. Carefully slide in the stoppers. The level rises. Read again. Thus :—

Weight of stoppers..... = 80·5 gm.
Volume of water, stoppers outside ... = 100 c.c.
,, ,, ,, ,, inside..... = 132·5 c.c.
∴ Volume of stoppers ..... = 32·5 c.c.
∴ Weight of an equal volume of water... = 32·5 gm.

Hence, relative density of the glass stoppers $= \frac{80{\cdot}5}{32{\cdot}5}$

$= \mathbf{2{\cdot}47.}$

With what kinds of solids could this method not be used?

EXERCISES

Using a 100 c.c. measuring jar, find the relative densities of—

(1) five glass marbles.
(2) fragments of slate pencil.
(3) lumps of marble.
(4) a copper cylinder.
(5) sand.
(6) a lump of aluminium.
(7) a cork.[1]
(8) a lump of beeswax, or of paraffin wax.

[1] Use a "sinker" which may be attached to the light solid by a piece of cotton. An old brass weight or a glass stopper will serve.

EXAMPLES IX

Relative Densities of Solids.

1. Find the relative densities of the following metals, given:—

(1) a litre of copper weighs 8950 grams.
(2) 75 c.c. of zinc weigh 540 grams.
(3) 676 c.c. of cork weigh 169 grams.
(4) 300 c c. of lead weigh 3·42 kilograms.

2. A rectangular block of marble measures 24 cm. long, 16 cm. broad, and 10 cm. in height, and it weighs 10·368 kilograms. Find its relative density.

3. A cube of wood measures 8 cm. along each edge, and its weight is 307·2 grams. Find its relative density.

4. Solids, weighing 10, 15, 20 grams respectively, were placed successively in a graduated cylinder partly filled with water, the surface of which read 95 c.c. originally. When the first was added, the water rose to 99 c.c., when the second was added, it rose to 106 c.c., and when the third was added to 108 c.c. Find the relative densities of the solids.

5. A brass chain weighs 96 grams. When lowered into a measuring jar already containing 100 c.c. of water, it raises the water level to 112 c.c. Find its relative density.

6. A dozen shot, made from lead, of relative density 11·4, raise the level of water in a measuring jar from 98 c.c. to 122 c.c. Find the weight of each shot.

7. A shilling weighs 87 grains, and has the relative density 10·4. What would be the weight of a platinum coin (21·5) of the same size?

8. Find the value of 1 c.c. of silver (relative density 10·5), when silver is worth 2*s.* 6*d.* per oz. Reckon 28 grams = 1 oz.

9. If 1000 oz. of water fill 1 cubic foot, and cast-iron is 7·2 times as heavy as water, find the weight of an iron block 4 in. long, 3 in. broad, and 1 in. high.

10. Find the volume of 1 cm. length of platinum wire (relative density 21·5), supposing a piece 1 metre long to weigh 2·15 grams.

(C)* THIRD METHOD: FRAGMENTARY SOLIDS

The volume of a solid cannot be determined to the tenth of a c.c. by means of a measuring jar. But by means of a balance we may weigh water to a centigram, that is, to the hundredth of

a cubic centimetre. Instead of noticing the rise of level in a measuring jar, we may find the loss of weight by overflow where fragments of a solid are put in a relative density flask full of water.

EXPT. 5. **To find the Relative Density of White Sand.**—Weigh a dry relative density flask. Introduce some sand, so that the flask is about one-third full. Weigh again. Nearly fill the flask with water, shake gently so as to dislodge air bubbles ; allow the sand to settle, and then carefully fill the flask with water, and introduce the stopper. Weigh again.

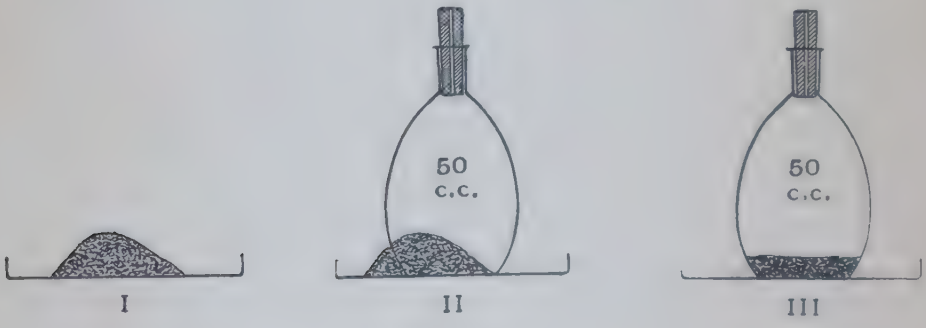

FIG. 29.

Lastly, wash the sand out of the flask, and weigh it filled with water alone.

Then calculate the weights represented in Fig. 29, I, II, and III.

Example :—

Weight of flask	=	14·80 gm.
Weight of flask + sand	=	23·84 ,,
Weight of flask + sand + water, *inside* ...	=	45·37 ,,
Weight of flask and water	=	39·80 ,,

Calculation :—

Weight of flask	=	14·80 ,,
Weight of flask + sand	=	23·84 ,,
∴ Weight of sand...	=	9·04 ... (i.)

Again—

Weight of flask + water	=	39·80 gm.
∴ Weight of flask + water + sand, *outside* ...	=	39·80 + 9·04
	=	48·84 gm. (ii.)

But—

Weight of flask + water + sand, *inside* ... = 45·37 gm. (iii.)

∴ Weight of water overflowing to make room for the sand... ... ... = 48·84 − 45·37 = 3·47 ,,

∴ Relative density of sand ... ... $= \frac{\text{Weight of sand}}{\text{Weight of overflow}}$

$= \frac{9{\cdot}04}{3{\cdot}47}$

$= \mathbf{2{\cdot}60}$

EXERCISES

Find the relative density of—

(1) powdered glass.
(2) lead shot.
(3) powdered antimony.
(4) powdered galena.

EXAMPLES X

Relative Densities of Solids

1. Find the relative density of sand from the following weights :— Flask, 25 grams. Flask + sand, 35 grams. Flask + sand, filled up with water, 56 grams. Flask filled with water, 50 grams.

2. Find the relative density of a powder from the following weights :—Flask, 25 grams. Flask filled with water, 50 grams. Flask + powder, 40 grams. Flask + powder filled up with water, 63 grams.

3. A 100 c.c. flask filled with water to the mark weighs 125 grams. When 57 grams of lead are added, and the water is again adjusted to the mark, it weighs 177 grams. Find the relative density of lead.

4. A 100 c.c. flask weighs 25 grams when empty. What will it weigh when 30 grams of powdered glass (relative density 3·0) are placed inside, and water is filled in up to the mark?

(D) FOURTH METHOD.

Principle of Archimedes.—There is another method of determining the relative densities of solids, which depends upon a fact discovered about 220 B.C. by the philosopher Archimedes

of Syracuse. It is said that Hiero, King of Syracuse, sent a lump of gold to a goldsmith's to be made into a crown. When the crown was delivered he suspected that, though of the right weight, it contained nevertheless an undue proportion of silver. So Archimedes was asked to find out the truth of the matter. Turning the matter over in his mind, it is said that he went to have a bath, and noticing the rise of the water he jumped out in great excitement, and ran naked through the streets of the city, crying, "Eureka, Eureka, I have discovered it!" But suppose we make the following experiment.

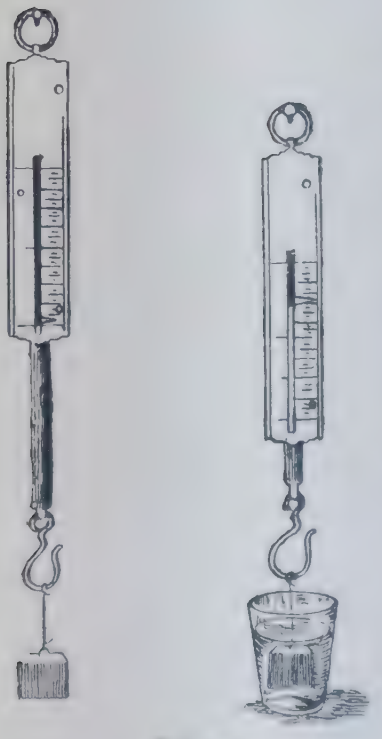

FIG. 30.

EXPT. 6. **To try whether a Solid weighs the same in Water as in Air.**—(1) Suspend a block of aluminium, or some other object, from a 100-gram spring balance, and read its weight in air. Then allow the block to sink below the surface of water in a tumbler, and again read the balance (Fig. 30). Is not the weight of the block in water much less than its weight in air? The water clearly supports part of the weight of the block.

Similarly, weigh both in air and in water an empty glass bottle and a block of wood. These have no weight in water, or rather the whole of their weight is supported by the water, and therefore they *float*.

(2) Let us next make an experiment which will show a connection between the loss in weight of a solid immersed in a liquid and the volume of the solid. Hang up a pair of hand scales. Take a brass cylinder which fits exactly into a bucket. Suspend the bucket under one pan of the scales, with the cylinder hanging underneath the bucket, and within an empty beaker as shown in Fig. 31, and weigh the whole. Then add water to the beaker until the cylinder is completely immersed. There is equilibrium no longer. Pour water carefully into the bucket, so as

to restore equilibrium, taking care that the cylinder is still completely immersed and not touching the beaker. How much water have you to add? At the moment when the bucket is filled up to the brim equilibrium is restored. So the cylinder when immersed in water lost a weight which was equal to the weight of water which occupied the same volume. The experiment might be repeated, using turpentine or a salt solution or some other liquid in place of water. This experiment has verified a fact which is usually called **The Principle of Archimedes**, viz.: **A solid when immersed in a liquid loses a weight which is equal to the weight of the liquid which would occupy the same volume.**

This is of great importance. For instance, it follows that if a block of aluminium, which in air weighs 95 grams, weighs only 60 grams in water, *i.e.* the loss of weight is 35 grams, the volume of the block must be 35 c.c. We can, in fact, readily find the volume of any solid which can be weighed in water.

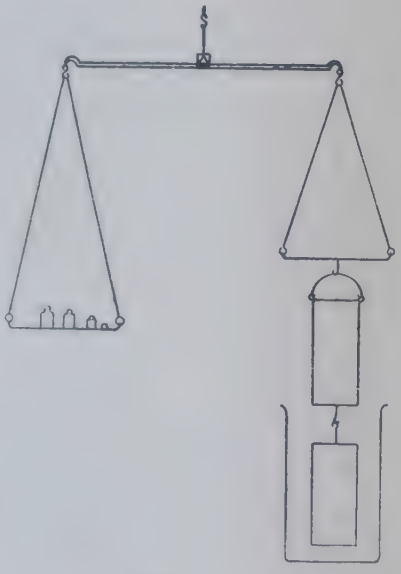

FIG. 31.

The Principle of Archimedes affords, therefore, a ready method for determining the relative density of any solid which can be suspended and weighed in water.

EXPT. 7. **To find the Relative Density of a Glass Stopper.**—Suspend a solid glass stopper by a thread above the left-hand pan of the balance (Fig. 32). Place a wooden bridge across the pan. Place a beaker on the bridge so that the stopper hangs freely within the beaker. Weigh the stopper. Pour water carefully into the beaker until the stopper is completely immersed, and remove any air bubbles which cling to the stopper with a camel's hair brush. Weigh it again. Thus in an experiment :—

Weight of glass stopper in air ... ... = 15 gm.
Weight of glass stopper in water ... ... = 9 ,,

Hence, loss in weight of glass stopper ... ... = 6 gm.
or, weight of an equal volume of water ... = 6 ,,

$$\therefore \text{ Relative density of glass stopper} = \frac{15}{6}$$
$$= 2{\cdot}5.$$

It is clear that the last *three* methods for determining the relative densities of solids cannot be applied directly in the case of one such as cork, which is lighter than water, or to another

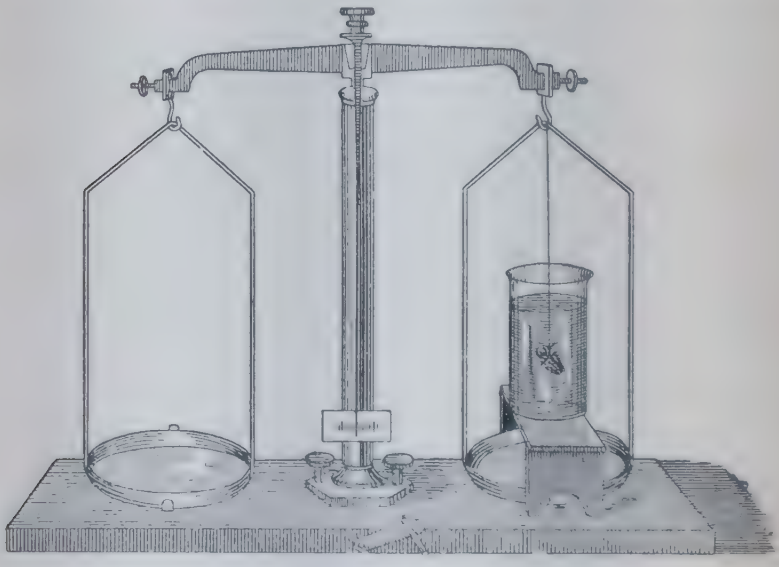

FIG. 32.

such as sugar, which will dissolve in water. The student is referred to other text-books for the special methods required in such cases.

Relative Densities of some Common Substances —

Cork	0·25	Tin	7·3
Oak	0·84	Zinc	7·2
Ice	0·93	Iron (cast)	7·2
White sand	2·6	Brass	8·4
Marble	2·7	Copper ...	8·9
Aluminium	2·7	Silver	10·5
Glass (crown) ...	2·5 to 3·1	Lead	11·4
Sulphur...	2·0	Gold	19·3
The Earth	5·5	Platinum	21·5
Antimony	6·7		

Exercises

Find the relative density of—

(1) a solid glass stopper.
(2) a lump of marble.
(3) a lump of sulphur.
(4) a copper cylinder.
(5) a brass cylinder.
(6) a lump of aluminium.
(7) an iron nut, screw, or bolt, or cylinder.

Examples XI

Relative Densities of Solids

1. A gold coin weighs 12·25 grams in air, and 11·55 grams in water. Find its relative density.

2. A lump of marble weighs 20·76 grams in air, and 13·14 grams in water. Find its relative density.

3. A lump of sulphur weighs 26·1 grams in air, and 13·07 grams in water. What is its relative density?

4. The volume of a glass stopper is 5·77 c.c., and its weight in air is 15·13 grams. Find its relative density.

5. A metal cylinder weighs 64·65 grams in air, and 57·4 grams in water. Of what metal does it probably consist?

6. If a silver coin weighs 41·6 grams, and has the relative density 10·4, what will it weigh in water?

7. Find the loss of weight in water of a 56 lb. cast-iron weight (relative density 7·2).

8. If a glass stopper weighs 24 grams in air, and 16 grams in spirits of relative density ·8, find the relative density of the glass.

9. A glass stopper weighs 30 grams in air, 17·5 grams in water, and 20 grams in spirits. Find the relative density of the spirits.

10. Find the relative density of a mixture of 100 c.c. lead (11·4) and 200 c.c. tin (7·3), assuming that they mix without contraction.

CHAPTER VII

THE MEASUREMENT OF THE PRESSURE OF THE AIR

The Existence of the Air.—We have no difficulty in realising the existence of solids and liquids around us. It is different with the air, and at first it is difficult to believe that there is an invisible substance surrounding us on every hand. Yet the existence of something invisible is at least suggested by the draught from an open window or door, or by waving corn and swaying branches.

EXPT. 1. **To show that an "Empty" Bottle contains Air.**—Take an empty bottle and force it mouth downwards into

FIG. 33.

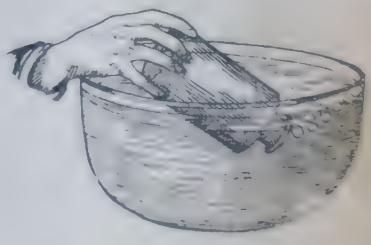

FIG. 34.

water. Notice that the water does not fill the bottle, showing that the bottle must contain something which opposes the entrance of the water. Then tilt the bottle, and notice that bubbles escape up through the water, and that only then does the water enter the bottle (Figs. 33 and 34).

Such facts help us to realise that we live in an ocean of air, which extends in all probability to a distance of at least 200 miles from the earth. This envelope of air around the earth is called the **atmosphere**.

We are aware of the pressure of the air in a strong wind, but in a room in which the air is at rest we are not conscious of any pressure.

EXPT. 2. **To show that Air is exerting a Pressure even when we cannot feel it.**—Bend a glass tube as in Fig. 35. Pour in some coloured water. The water stands at exactly the same height in both arms. Why? If we attach a piece of rubber-tubing to the arm A B, and suck out some of the air from that arm, what may occur? Try. The water rises in the arm A B, and falls in the other, as represented in Fig. 36. Secondly, blow in a little air. The water falls in the one arm, and rises in the other.

The rise and fall of the water in the arm A B shows that the air was originally exerting a pressure equally upon the surface of the liquid in both arms and that the pressure of the air in A B was diminished when air was sucked out, and increased when air was blown into it.

FIG. 35.

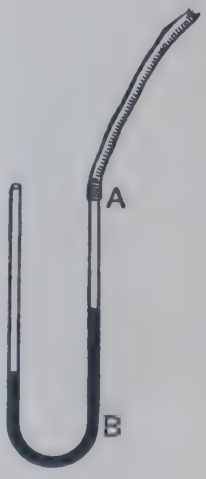

FIG. 36.

EXPT. 3.—Place one end of a *pipette* below the surface of some water. Apply the mouth at the other end, and suck gently. As the air is gradually inhaled, the water rises in the pipette.

This suggests that the air is exerting a pressure on the surface of the water outside the pipette, so that the water is forced up the tube as soon as the air is partially removed from the inside, and when the pressure there is consequently diminished.

The action of a *syringe* or that of a *suction-pump* can now be understood to depend upon the pressure which the air exerts upon the surface of a liquid.

EXPT. 4.—Take a tumbler, or, better still, a gas jar with a ground edge, and plunge it into water so that it is completely

filled. Press a piece of cardboard against the mouth. Carefully lift the vessel out of the water, mouth downwards. Note that the cardboard remains over the mouth. Why? Clearly it is held there by some pressure greater than the weight of the water within the tumbler (see Fig. 37).

EXPT. 5.—Fit a thistle-funnel in one end of a piece of thick rubber pump-tubing. Fix the other end to an air-pump. Moisten the palm of the hand, and press the mouth of the funnel against it. Then proceed to exhaust the air from the tubing. Observe that the funnel is pressed firmly against the hand, so that some force is necessary in order to remove it.

It is clear, therefore, that the air even when quite still is exerting pressure, and the fact that no force may be felt upon the extended hand arises in the first place from the pressure being exerted above and below it equally. Then why is the hand not squeezed flat? Because there is an internal pressure of the blood in the veins and arteries and other fluids within the body, greater than that of the atmosphere.

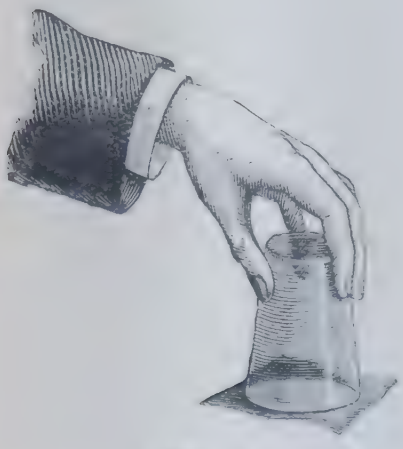

FIG. 37.

Repeat Expt. 5, applying the mouth of the funnel to the cheek. As soon as the pressure of the air within the funnel is diminished, the pressure within the cheek causes the flesh to swell outwards, while the funnel is pressed by the air outside against the cheek.

The Pressure of the Air is due to the Fact that Air has Weight.—We are not conscious of the air having any weight, but by the following experiment we can readily show that it can be weighed.

EXPT. 6. **To show that the Air has Weight.**—Take a round bottomed flask and fit it with a one-holed rubber stopper through which is passed a short glass tube fitted with a piece of rubber tubing and a clip, or, better, with a glass tap (Fig. 38). Suspend the flask by a string from one arm of a balance, and

weigh it. Then, leaving the weights on the pan, remove the flask, and exhaust some of the air by means of an air-syringe. Again suspend the flask on the balance, and observe that it is now lighter than the weights. Lastly open the tap : air is heard rushing into the flask, and it is then found to have regained its original weight.

It is clear therefore that the air has weight. As a matter of fact 1 litre of air weighs only 1·2 gram. But just as the weight of one eider-down quilt may be scarcely felt, yet many of them one above the other would be very oppressive, so, extending as the air does to a distance of at least 200 miles from the surface of the earth, it may be understood that the very considerable pressure of the air at the surface of the earth is due to the fact that the air has weight, slight though this be.

FIG. 38.

Torricelli, an Italian physicist, first showed in 1643 in the following way how to measure the pressure which the air is always exerting.

EXPT. 7. **To measure the Pressure of the Air.**—Take a thick-walled glass tube about 4 mm. internal diameter, and 90 cm. in length. Seal one end in the blow-pipe flame. Holding the open end upwards, almost fill the tube by pouring clean dry mercury[1] into it through a small funnel. Gently tap the tube so as to dislodge air bubbles. Pass a bubble the length of the tube so as to sweep out smaller bubbles. Fill the tube completely with mercury. Close the open end firmly with the thumb, and invert the tube in a mortar containing mercury (Fig. 39). The mercury sinks in the tube, leaving a clear space A B at the top of the tube. Tilt the tube : if the mercury does not rise completely to the top, some air bubbles have not been removed ; these must then be displaced by inverting the tube again, and gently tapping. Measure the length of the mercurial column above the level of the mercury in the mortar. It is about 76 cm., or 30 inches. From the way in which the tube

[1] On Management of Mercury, see Appendix, Vol. II.

was set up it seems impossible that the space A B can contain any air.

Let us repeat Expt. 7 in another form.

EXPT. 8.—Take a thick-walled glass tube about 90 cm. long, closed at one end, and fix a piece of thick rubber tubing about 6 inches long upon the open end. Fix a short glass tube open at both ends into the free end of the rubber tube. Resting the closed end upon the bench, carefully pour mercury into the apparatus until it reaches half way up the short tube. Sweep out air bubbles as before. Then fix the apparatus upon a board as shown in Fig. 40.

FIG. 39.

The level of the mercury in the closed arm is again about 30 inches above that in the open arm. If Expts. 7 and 8 were repeated with wider or longer tubes, the length of the mercurial column would still be found to be about 30 inches or 760 mm.

EXPT. 9. **To show that A B is a Vacuum.**—(1) Tilt the tube used in Expt. 7 gradually, keeping the unsealed end still below the mercury in the mortar. The mercury rises in the tube, and finally completely fills it (Fig. 41).

(2) Pour some water above the mercury in the mortar. Raise the tube carefully so that the open end C lies in the water. The mercury at once runs down, and water not only takes its place, but completely fills the tube.

It is evident, therefore, that the space A B is free from air. After the author of the experiment this space is usually called the **Torricellian vacuum**.

From this and the former experiments it appears that the *column of mercury B C must be supported by the pressure of the air on the mercury in the mortar.*

FIG. 40.

So also it appears that in Expt. 8 the column of mercury in the closed arm must be supported by the pressure of the air on the mercury in the short open arm.

EXPT. 10. **To show that the Length of the Mercurial Column varies with the Pressure of the Air.**—(1) Connect an exhausting air-syringe by thick rubber tubing to the short open arm of the apparatus used in Expt. 8. Exhaust some of the air from the short arm : the mercury immediately falls several inches in the other arm. Readmit air : the mercury rises to its former level.

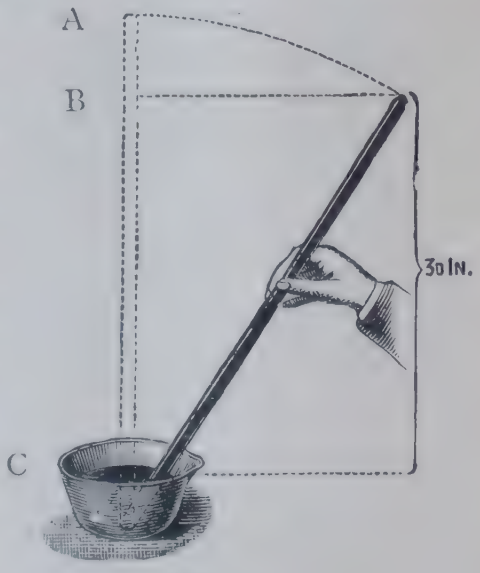

FIG. 41.

(2) Connect an air condenser in the same way with the open end. Gently force a little air into the short arm : the mercury is forced right up to the top of the closed arm.

Robert Boyle, who invented the air-pump, placed the whole apparatus used in Expt. 7 within the receiver of the air-pump. He found to his delight that the mercury fell further and further in the tube as the air was gradually exhausted, until it was only an inch or two above the level of the mercury outside. On re-admitting air into the receiver, he found that the mercury rose to its original level in the tube.

It is evident, therefore, that the **height of the mercury in**

the tube above the level outside is a measure of the pressure of the air. The instrument was called by Robert Boyle a **barometer,**[1] and ever since his time it has been in use for this purpose.

Variations in the Atmospheric Pressure.—If we imagine the earth to be enclosed in a vast ocean of air, we should expect the pressure of the air to become less and less as we ascend above sea-level, and this should be made evident by our barometer. It was shown to be so by Périer in 1648. He ascended the Puy-de-Dôme, in the Auvergne Mountains of France, and found that the mercury fell about 3 inches. The barometer even shows that in the course of any railway journey there are variations in the atmospheric pressure according to the height above sea-level.

Further, at any one place, for example at Manchester, the barometer shows that variations occur from time to time in the atmospheric pressure. These variations arise from the movements which are continually taking place in our atmosphere, including changes in the amount of moisture present in the air. Other things being equal, moist air is not so heavy as dry air: this may appear strange, and it will probably not be understood till later. Sometimes in a storm the height of the barometer will change by as much as 1 inch within 24 hours.

In the West Indies there are frequently very violent storms, during which the height of the barometer will vary far more rapidly than is the case in the British Isles. For instance, at St. Vincent on September 8, 1898, the barometer stood at 10 a.m. at 29·54 inches, and the wind was blowing strongly from the north-east. The wind then veered rapidly round, and by 11 a.m. a hurricane was blowing from the west, powerful enough to uproot the largest trees, and at 11.40 the barometer had fallen in 100 minutes more than 1 inch, down to 28·51 inches. The wind then died away, and there was a dead calm for three-quarters of an hour, and during this time the barometer was quite steady. At 12.25 p.m. the wind began again from the south, and by 3.0 p.m. the barometer had risen to 29·53 inches. During the day it also rained in torrents: between 9 a.m. and noon actually 4·9 inches of rain fell.

Note.—At this point, pp. 170–7 of Chapter XVIII. may be

[1] Greek: *baros*, weight, and *metron*, a measure.

taken with advantage, showing how to represent in a diagram a number of readings of a barometer.

A Weather Glass.—The weather glass shown in Fig. 42A, which is frequently seen hung up in houses, contains a barometer of the kind shown in Fig. 40. On the mercury in the short arm rests a little float, which will rise or fall as the level of the mercury alters. This float is connected by a string which passes

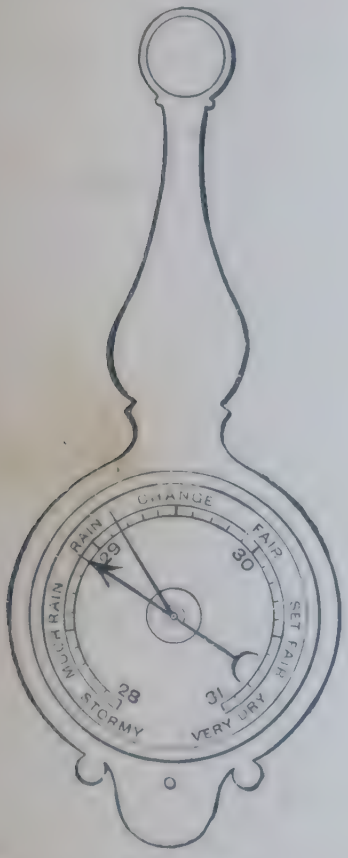

FIG. 42A.

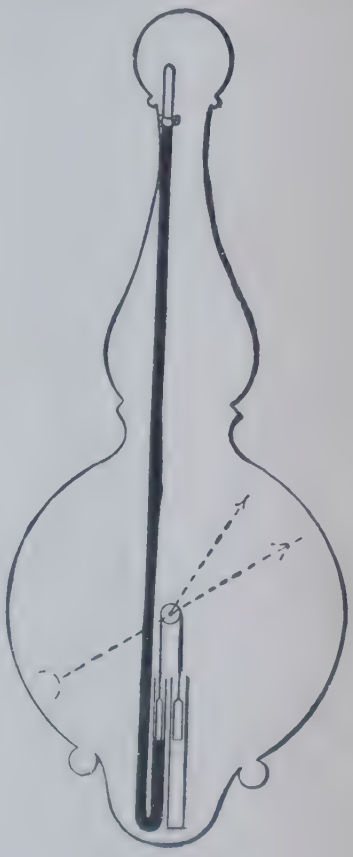

FIG. 42B.

over a wheel with a weight, which is slightly lighter than itself. The apparatus, as seen from the back, is shown in Fig. 42B. To the wheel is fixed a long needle, which is to be seen on the face of the apparatus. The motion of the float will clearly cause the wheel and also the needle to turn round. Numbers are engraved upon the face of the circle round which the needle can move, and the number to which the needle points shows the height of

the mercury in the tube. The words *rain*, *change*, *fair*, etc., are also engraved upon the circle.

The Standard Atmospheric Pressure.—The *average* pressure, however, at any place on the sea-level is represented by about **30 inches** or **760 mm. of mercury.** It is customary, moreover, to regard the pressure which is measured by 760 mm. of mercury as the standard atmospheric pressure. In other words, the pressure exerted by a column of mercury 760 mm. in length is called **a pressure of one atmosphere.**

FIG. 43A.

FIG. 43B.

It is useful to remember that this pressure is about 14·7 lbs. weight upon every square inch, or 1033·6 grams weight upon every square centimetre.

A Water Barometer.—How long a column of water could be supported by the pressure of the air? We must find out how long a column of water would exert the same pressure as a column of mercury 30 inches in length.

EXPT. 11. **To find out how many inches of Water will balance 1 inch of Mercury.**—Bend a long glass tube into the form of a **U**, and fix it vertically. Pour a little mercury down one arm, so as to fill the bend and rise about 2 inches up the arms; it stands at the same level in both arms (Fig. 43A). Then incline the tube, and introduce water by means of a pipette into one arm, so that the column of water is between 13 and 14 inches in length. Fix the tube vertically again. Measure CD, the length of the water column (Fig. 43B); it is, say, 13·6 inches. What balances this column of water? The short column of mercury FE *above* the level of the mercury in the other arm. Measure FE: it is 1 inch.

Thus a column of mercury 1 inch in length exerts the same pressure as that produced by a column of water 13·6 inches long.

Hence, a column of mercury 30 inches long will exert the same pressure as that produced by a column of water $\frac{13{\cdot}6 \times 30}{12}$ or 34 feet long. The tube of a water barometer must, therefore, be more than 34 feet in length.

Glycerine has occasionally been used as the liquid in a barometer: the tube must be more than 27 feet in length.

It is clear that in a water or glycerine barometer variations in height caused by the changes in the atmospheric pressure will be much greater and more evident than in a mercury barometer.

Exercises

1. Attach a U-tube containing water by means of a rubber tube to the coal-gas supply. What do you observe? Can the pressure of the gas be measured?

2. Find how many inches of glycerine will balance 1 inch of mercury.

CHAPTER VIII

FITTING UP APPARATUS

NOTHING is more frequently needed in any laboratory[1] than some source of heat. If a chemist had no fire or flame at his service, he would be able to do very little. The alchemists generally used some form of furnace. In quite modern times[2] coal gas has come into common use, and is supplied in every laboratory. Is an ordinary bright coal gas flame suitable for laboratory purposes?

Hold a piece of white cardboard for a few moments across a bright gas flame. It becomes very sooty. If, then, a glass vessel were heated in such a flame, it would quickly become coated with soot, and it would be difficult to watch the progress of what might be going on inside it. But it is frequently most important to be able to see the contents of a vessel. Consequently, a bright coal gas flame is not a satisfactory source of heat, and it would be a very great gain to have a flame which would not deposit a black curtain of soot upon any vessel heated by it. A burner which would give such a flame was devised by Robert Wilhelm Bunsen, Professor of Chemistry in the University of Heidelberg. A great teacher and a famous investigator, he died in 1899, at the ripe age of 88 years.

The Bunsen Burner.—If air be allowed to mix with coal gas before it is burnt, a blue flame hotter and smaller than the ordinary luminous gas flame results. Examine a Bunsen burner. At the bottom of the tube A (Fig. 44) is a collar B. Oval windows or holes are pierced in the tube and in the collar

[1] See Appendix on Laboratory Fittings, and Lists of Apparatus and Chemicals.

[2] The streets of London were first lighted with coal gas in 1812.

Either the tube or the collar can rotate so that these windows can be set opposite to each other, admitting air or alternately excluding air.

(1) Close the holes. Light the gas. The flame is yellow. Bring a piece of white cardboard into the flame; it becomes sooty.

(2) Open the holes. The flame becomes blue. Hold a piece of cardboard in it. It becomes brown and singed, but not sooty. Hold a match across the flame about one inch above the tube. The wood is burnt in two places, showing that there is an inner cooler region. Try whether you can introduce a match head into the middle of the flame without lighting it.

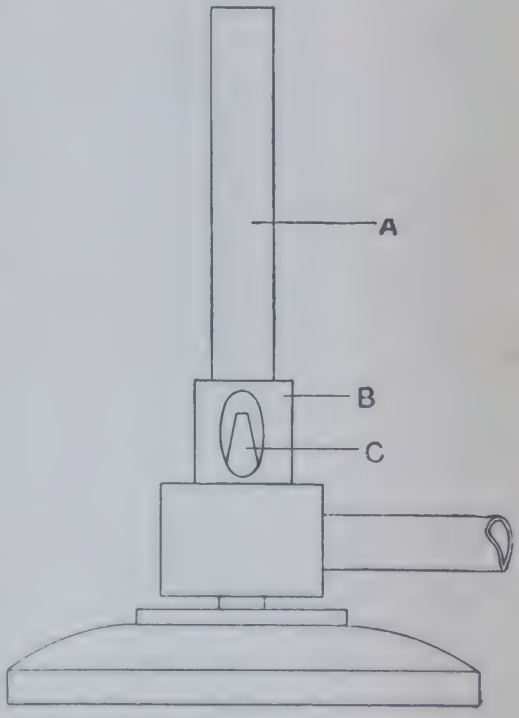

FIG. 44.

Is the blue flame hotter than the yellow flame? Yes. Does closing the windows make any difference to the steadiness of the flame? Yes, the yellow flame is taller, and much less steady.

(3) Light the burner with a blue flame. The tube contains a mixture of gas and air. Press gently upon the rubber gas supply tube, making the flame smaller and smaller. Suddenly the flame will shoot down the tube, and it will be seen *burning below*. Notice the curious smell. Light the burner at the top. The flame is yellow, although the holes are open. Any attempt to turn the tube will now lead to burnt fingers, since the inner flame makes the tube very hot. It is better to turn the gas out, and then light it again.

When a very small flame is required, the supply of air should be reduced as well as the coal gas.

Glass is useful to the chemist on account of (1) its trans-

parency, (2) the ease with which it can be worked into any shape when hot and plastic, (3) its not being attacked by most chemicals. Its disadvantage is its brittleness. It is particularly liable to break on heating, because it is a bad conductor of heat, and therefore one part of the glass expands on being heated before neighbouring parts become warm.

To Illustrate the Liability of Glass to Crack when Heated.—(1) Heat one end of a thick, hard glass tube in the flame of the foot-blowpipe. The glass cracks before the fingers feel the heat.

(2) Warm another piece of glass tube very gradually. Then pour cold water on it. It cracks.

(3) Fill a test-tube one-third full of water. Hold it in the flame so that the flame plays on the glass near the water line. The tube cracks.

To Boil Water in a Glass Vessel.—(1) Try again to warm water in a test-tube, as follows. Bend a doubled strip of paper back on itself round the upper part of the tube, and hold the tube by the paper. Let the flame play on the bottom of the test-tube. Shake the tube gently so as to keep the water dancing up and down the sides of the tube, and so warm them gradually. The water boils without accident.

FIG. 45.

(2) Place a piece of iron *wire gauze* on a tripod-stand, or on a ring of a retort stand. Place a glass beaker or flask containing 50 c.c. water on the gauze. Put a Bunsen flame under the gauze. The gauze distributes the heat over the bottom of the vessel, and the water can be heated until it boils without accident.

(3) Place a little sand in an iron tray, and put the tray upon a tripod-stand (Fig. 45). Place a beaker containing 50 c.c. water upon the sand, and heat the tray with a Bunsen burner. The heat is distributed by the sand, so that the beaker can be heated safely until the water boils. Observe that the water heated over wire gauze boils in much less time than that heated on the sand.

The tray containing the sand is called a *sand-bath*.

To Make a Wash-Bottle

It will generally be possible to fit up for ourselves the apparatus required for our experiments. We shall require glass flasks, corks, glass tubing, and india-rubber tubing, and some knowledge of how to use them.

Suppose we learn in the first place how to make the apparatus shown in Fig. 46. This is called a *wash-bottle.* It is useful for sending a stream of water in any desired direction. By blowing at *d* a jet of water spurts from *c*, and the india-rubber connection *e* allows of the water being sent in any direction. We shall require an 8-oz. flask, a cork, about 50 cm. of glass tubing, and a small piece of rubber tubing.

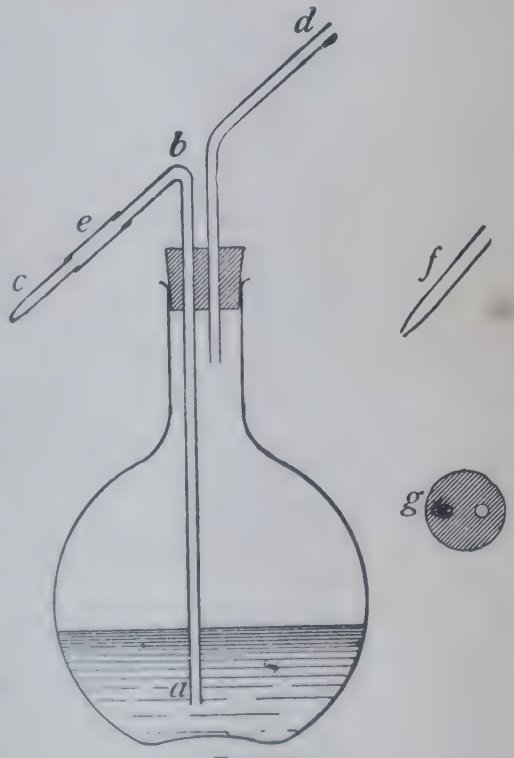

FIG. 46.

The cork (wooden) selected must be just too large to fit into the mouth of the flask. It can be made to fit the neck by rolling and squeezing it under the foot. Begin gently, turning it round and round, and press harder and harder. The cork now fits well into the neck of the flask, and being somewhat elastic presses tightly against the glass all round.

An india-rubber stopper may be used instead. It should fit the mouth easily, since it will be enlarged when the glass tubes are passed through it.

To bore a Cork.—(1) A hollow brass tube sharpened at one end may be used, called a cork-borer. Choose one a little narrower than the glass tubes which are to go through the cork. At the narrow end of the cork put marks where the holes are to be made. Stand the cork on the table, narrow end up. Keep the

borer carefully upright. Drive it gently into the cork with a screw motion, half a turn at a time, pausing to see that it keeps upright. Along with a set of cork-borers is supplied a little solid brass rod, which can be used both to press the core of cork out of the borer and as a handle to turn the borer round. (The borers will require sharpening periodically.)

(2) Or, heat the end of a rat-tail file in a flame. Drive it gently through the cork. Enlarge the hole with the file, but take care that this is done evenly, so that the hole remains round. It should not be made so large that the tube fits loosely into the hole.

To cut Glass Tubes.—Take a sharp triangular file. Make a nick in the glass tube about 15 cm. from one end. Hold the tube in both hands. The two forefingers should be together

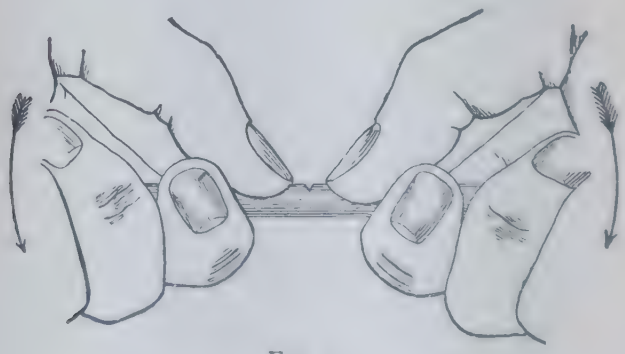

FIG. 47.

immediately underneath the nick. The two thumbs are above. Turn the wrists outwards and downwards, round the forefingers as axes, at the same time pulling the hands apart (Fig. 47). The glass cracks clean across.

To bend Glass Tubing.—Hold a piece of glass tubing lengthwise along a *flat* gas flame, as shown in Fig. 48 (fish-tail or batswing, not the Bunsen flame). Keep it moving slowly round and round, and also to and fro. As soon as it feels slightly soft, hold it by one hand alone, and cease rotating it. The free end will gradually fall (Fig. 49). As soon as it is bent to the required angle, remove it from the flame. A considerable length has been evenly heated, and a smooth bend has resulted. Now cut two lengths of 15 cm. and 25 cm., and bend them at angles of 135° and 45° respectively, as shown in Fig. 46.

Mistakes to avoid in bending Glass.—(1) Hold a piece of glass tubing across the blue Bunsen flame. It becomes hot

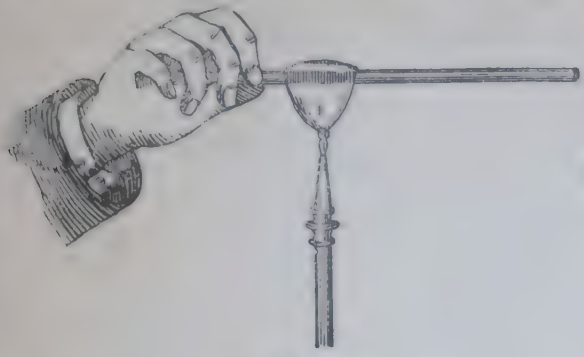

FIG. 48.

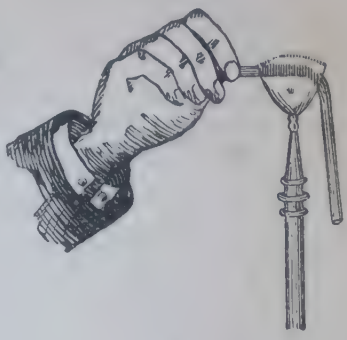

FIG. 49.

in two places. When soft, bend it. The glass is found bent in two places. (2) Hold a piece of glass tubing across an ordinary flat gas flame. The tube is heated this time in one place. Bend it. The tube is flattened on the outside of the bend and thickened on the inside. When cold, try to bend the tube. It snaps easily at the bend.

To round the Ends of Glass Tubing.—The sharp edges left at the ends of the glass tubing must be rounded, otherwise the tubes will not slide easily through corks, and are apt to cut any rubber tubes that are slipped over them. Round the ends of the tubes just made—(*a*) by grinding the sharp edge on a file or on a stone floor, (*b*) by holding the end in the edge of the blue flame, at a point about two-thirds up the flame, as shown in Fig. 53, until the glass just melts at its edges.

To draw out Glass Tubing.—

(1) Holding a piece of glass tube in both hands, heat it at one place in the *blue Bunsen* flame. Remove it when soft from the flame and pull.

The tube is drawn a little at A (Fig. 50). At B it is pulled

FIG. 50.

out much finer, because it happens to be hotter, and it has broken there because the pull was too hard. The lump C was

in the cooler central portion of the flame, and never became properly heated.

(2). Heat another tube in a *flat flame*, holding it lengthwise and turning it slowly round and round. Pull the two ends apart. There is no lump, because the tube was heated evenly,

FIG. 51.

but the finely drawn out tube has softened at once, and is bent down (Fig. 51).

(3) Heat another tube as in Expt. 2 until the tube feels limp in the fingers. *Remove it from the flame.* Draw slowly at first, then more quickly, but gently rotating each end until the hands are wide apart. Hold it rigidly for a few seconds, until the glass stiffens (Fig. 52). The glass is now drawn out perhaps a yard long or more. Cut the tube with a very sharp file gently at A and B. Notice how elastic this section A B is. It can be bent nearly to a circle. If bent too far, it snaps, and the pieces spring back straight again. Dip the end B into water, and blow at A. Bubbles arise. Dip another piece of this fine tube into an inkpot, and note the result. Then, A B is

FIG. 52.

really hollow all through. The end A C can be used for the wash-bottle jet. The very narrow tube A B is called "capillary tubing" (Lat. *capilla*, a hair).

The foot blowpipe flame is better than a flat flame for drawing out glass tubing.

To narrow or close an End of a Glass Tube. – Incline a glass tube 20 cm. long downwards at an angle of 60°, and hold the lower end in the edge of a Bunsen flame at a point about two-thirds of the height of the flame, as shown in Fig. 53. The edge becomes red-hot, fuses, and gradually contracts.

Remove the tube from the flame before it is quite sealed up. Cut the glass about 4 cm. from the end. This portion can be used for the wash-bottle jet.

Hold one end of the remainder of the tube for a longer time in the flame, and close the end entirely.

Rubber tubing sometimes becomes stiff when not in use. Warm gently and pull it about; it soon softens. Cut off a piece an inch and a half long for future use. If a rubber tube is too wide for a glass tube, it can be made to fit by doubling it back on itself.

To fit together the Wash-Bottle.—Moisten the end *a* of the long tube *a b* (Fig. 46). Push the long tube through the cork with a zig-zag screw motion. Hold the tube near the cork, and avoid holding it at the bend *b*, where it might break and cause a nasty cut in the palm of the hand. Fit *d* in the cork also. Now fit the cork into the bottle. The end *a* should come nearly to the bottom of the flask, but not so near as to risk touching it.

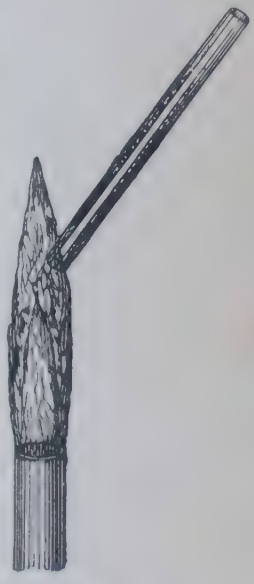

FIG. 53.

NOTE.—Tubes should never be pushed into corks when they are fixed in the mouths of flasks; nor should corks be fixed in the necks of flasks when the flasks are standing on the bench; hold the flask in the left hand.

To test whether the Flask is Air-tight.—Close the glass tube at *e* and blow at *d*. Does the air hiss out through the cork? Again closing *e* suck at *d*. The tongue is sucked back, and held at the end of the tube. Attach the jet *c* by means of the rubber tubing. If the flask is now filled with water, it is ready for use.

To use the Wash-Bottle.—

(1) Blow at *d*. A fine jet of water issues from *c*.

(2) Invert the bottle. A large stream of water issues from *d*. It is often well to keep the bottle filled with distilled water. If a supply of hot water is required, the bottle can be heated over a flame upon wire gauze.

LABORATORY RULES, if not provided by a teacher, should be made by the student for himself :—

1. Each boy must use his own apparatus and work at his own place. Experiments, calculations, or notes must not be made at the bench reserved for balances.
2. Clean apparatus, and put it away in its proper place as soon as done with.
3. Bench, drawers, and cupboards must be left clean and in order. Test-tubes must be left clean, and half full of water.
4. Solids (*e.g.* matches, filter-paper, rubbish) must not be thrown into the sink, but be put into the refuse pan.
5. Never pour strong acids down the sink. Dilute them with water before throwing them away.
6. Stoppers must not be left out of their bottles on the bench, but be replaced.

EXERCISES

Working of Glass. Fitting up Apparatus

1. Cut a tube 10 cm. in length. Round the edges. Bore a cork to fit the tube tightly. Pass the tube through the cork.

FIG. 54

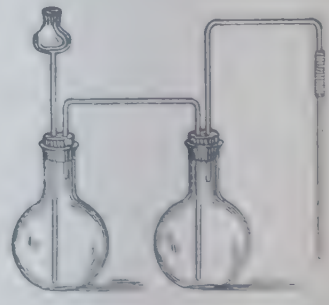

FIG. 55.

2. Prepare tubes as follows :—

a tube 15 cm. long bent at 120° in the middle.
" 15 " " " " 90° " " "
" 15 " " " " 60° " " "
" 40 " " " into an equilateral triangle.

3. Cut a tube 40 cm. long. Bend it at right angles near one end, and at 60° near the other end, taking care that both bends are in the same plane.

4. Cut two tubes 10 cm. long. Round their edges. Bore two holes through a cork to fit the tubes. Fix the tubes through the cork.

5. Cut a tube 25 cm. long. Draw it out near one end, and cut it

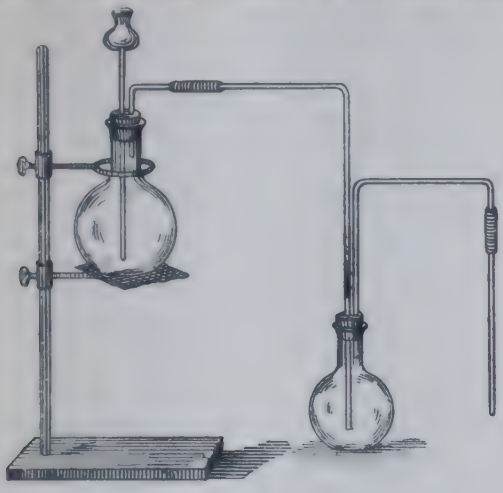

FIG. 56.

in two. The shorter pieces can serve as wash-bottle jets, the longer ones can be used as pipettes.

6. Practise pulling out glass tubes to full arms' length. Break up the capillary tube so formed into 6-inch lengths, and save them.

7. Fit up the pieces of apparatus shown in Figs. 54, 55, and 56.

CHAPTER IX

CHANGE OF STATE—LIQUEFACTION AND SOLIDIFICATION

We are acquainted with many different solids and liquids.

What is the effect of heat upon them?

Many solids are changed into liquids by heat, which solidify again on cooling.

EXPT. 1.—Place some powdered **sulphur** in a test-tube. Heat very gently in a Bunsen flame. The sulphur quickly melts, and forms a yellow liquid. The liquid becomes thick like treacle, and of a dark red colour. Then it becomes more liquid again, and begins to boil. Let it cool. The liquid quickly solidifies again, forming a solid cake of sulphur.

EXPT. 2.—Place a lump of **paraffin wax** in a test-tube, and heat it gently in a flame. The wax melts very quickly, and forms a clear liquid. Let it cool. The liquid is quickly changed into an opaque white solid mass.

EXPT. 3.—Place a few pieces of **tin** in an iron spoon. Support it on a tripod. Heat in a good Bunsen flame. The tin melts completely. Skim off the earth-like dross on the surface of the molten metal with an iron file. Pour the liquid in a thin stream into a mortar full of water. It is converted into a bright crumpled mass of the solid metal again.

EXERCISES

Describe the effect of heat and subsequent cooling upon the following substances. In each case describe the appearance of the substance before heating it.

(1) sealing wax.
(2) shellac.
(3) vaseline.
(4) bees' wax.
(5) material of a stearin candle.
(6) lead.
(7) zinc.
(8) copper.
(9) common salt.

A further examination of the effect of heating metals in the air and of the dross that appears upon the surface of molten metals will be made in Volume II. Chapter II.

Let us study one particular case more fully.

EXPT. 4. **The Effect of Heat upon Ice.**—Break some ice into small fragments, the smaller the better. Pack them into the lower half of a wide beaker. Plunge a thermometer into the mass. The temperature is 0° C. Warm the beaker on wire gauze over a very small flame, and stir the mass continually with the thermometer, occasionally removing the flame for a few seconds. Read the thermometer frequently. The ice rapidly melts. Although so much heat is added, nevertheless the temperature remains fixed at 0° to 1° C., until the whole of the ice is melted. Only then does the temperature begin to rise, and continue to do so.

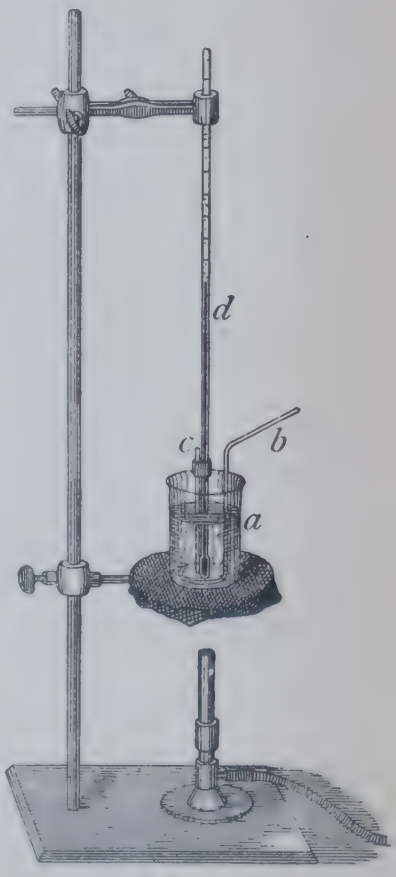

FIG. 57.

It is evident that while ice is being liquefied *no rise* of temperature occurs. In other words, the change of state takes place at a definite temperature. This temperature is called the **melting point** of the ice.

Other substances behave similarly to ice, and we may now proceed to determine in a simple way the melting points of some of them.

*EXPT. 5. **To find the Melting Point of Paraffin Wax.**—Take a narrow glass tube. Close one end by holding it in a flame. Introduce some small particles of paraffin wax to a depth of about 0·5 cm. Attach it to a thermometer by means of a thin slice of rubber tubing so that the wax is opposite the middle of the bulb. Support a beaker, of about 100 c.c. capacity, half-filled with water, over wire

gauze on a ring of a retort-stand. Place a glass stirrer, *b*, inside, the end of which has been bent into a circle at right-angles to the stem. Support the thermometer so that its bulb is in the centre of the water (Fig. 57). Heat gently, stirring constantly. As soon as the paraffin wax melts, read the temperature. Probably you have been too late, and the reading is too high.

Allow the water to cool slightly. Take a fresh capillary tube containing paraffin wax and ascertain the *exact* melting point, removing the flame from time to time, and stirring thoroughly, so that the temperature rises very slowly. Read to the nearest degree.

Exercises[1]

Find the melting points of (1) iodine, (2) naphthalene, (3) washing-soda, (4) Glauber's salt, (5) sulphur, (6) "hypo," (7) phenol.

We have found that different solids melt at different temperatures. Usually, moreover, the melting point is in each case very definite. Pure sulphur, for example, has a well-defined melting point. Impure substances, on the other hand, do not melt entirely at one temperature, but gradually through a range of temperature. Hence, use may be made of melting points in order to distinguish and identify substances, and to detect the admixture of foreign ingredients. The melting point is indeed one of the best defined of physical "constants." Here are a few melting points.

Table of Melting Points

Platinum	1770° C.	Naphthalene ...	79° C.
Gold	1062°	Paraffin, about...	54°
Copper	1054°	"Hypo"	48°
Silver	961°	Washing-soda ...	34°
Common salt ...	851°	Butter, about ...	33°
Zinc	419°	Glauber's salt ...	32°
Lead	326°	Ice	0°
Tin	233°	Mercury	−39°
Sulphur	115°		

[1] It will be found necessary to use strong sulphuric acid in place of water in the case of iodine and sulphur.

Latent Heat of Liquefaction.—We found that while ice was being melted no rise of temperature occurred, in spite of the addition of so much heat. In a similar way heat apparently disappears while any other solid is being changed into a liquid. This heat which alters the state of a solid, but does not raise its temperature, and which therefore apparently disappears, was first investigated by Dr. Black of Edinburgh about the year 1760. The question may be asked, whether, when such a liquid is reconverted into the solid state, the heat will become sensible again.

EXPT. 6. **To find whether Heat is given out when a Liquid solidifies.**—Take some crystals of a substance known to the photographer as "hypo," and to the chemist as sodium thiosulphate. Put them in a perfectly clean flask. Heat gently, until the crystals are completely melted. Introduce a thermometer. Place a plug of cotton-wool in the neck of the flask. Allow to cool in a quiet place. Then remove the plug, and introduce a small crystal of sodium thiosulphate. The liquid quickly solidifies. At the same time the thermometer rises many degrees, and the flask becomes quite hot.

It is evident then that the heat, which apparently disappears when a solid is liquefied, can be recovered if the liquid be again caused to solidify. Hence the heat is said to be hidden or latent, and it is usually spoken of as **latent heat of liquefaction.**

Freezing Mixtures.—Advantage is taken of the fact that heat becomes latent in the liquefaction of any substance to produce "freezing mixtures." For, supposing the liquefaction of a substance is induced to take place, and yet no heat is supplied from any external source, the heat which must be absorbed in the process of liquefaction will be taken out of the substance itself or out of anything in contact with it, and a considerable fall of temperature may ensue. For instance, if 1 part of common salt is mixed intimately with 4 parts of pounded ice or snow, the mass liquefies, because the freezing point of *brine* is far below that of water. Hence, at the temperature of the experiment, the mixed materials ought to be liquid. But in order to liquefy they must absorb heat. They, therefore, sacrifice their own temperature in order to get this heat, and the temperature sinks to about -22° C.

Problems

Find whether freezing mixtures can be made from—

(1) Glauber's salt and strong hydrochloric acid.
(2) Ice and crystallised chloride of calcium.
(3) Salt and water.
(4) Sal-ammoniac and water.

CHAPTER X

CHANGE OF STATE—VAPORISATION AND CONDENSATION

We have found that many solids are changed into liquids by heat. What is the effect of heat upon liquids?

Many liquids are changed into vapours by heat, which liquefy again or solidify on cooling.

Expt. 1.—Place a few crystals of **iodine** in a test-tube. Heat the end of the tube very cautiously, holding the tube nearly level. Some of the iodine melts, but at once heavy violet vapours are given off. A dark deposit forms on the upper and cooler part of the tube. Examine it carefully. Crystals of iodine have been formed. The violet vapour of iodine when cooled has become solid iodine again. The vapour is said to be **condensed.** Finally, all the iodine disappears from the bottom of the tube

Expt. 2.—Place a drop of **mercury** in a test-tube. Heat it carefully in a flame, holding the tube nearly level. A mirror of minute globules of mercury is soon seen upon the upper and cooler portion of the tube, although there is no visible vapour. Then the mercury boils, and the globules grow larger as the hot invisible vapour is cooled and condensed, and some of them may roll down and again be vaporised

Expt. 3.—Heat some **water** in a flask over wire gauze. A rolling vapour is soon seen within the flask, and a dew is formed on the cool sides of the flask. Bubbles have appeared on the bottom of the flask, and some small ones may ascend through the water. The bubbles grow larger, and more and more reach the surface of the water, which is at last completely broken up, and a white cloud of "*steam*" is seen at the mouth of the flask. The water is then said to *boil*. Nothing is now seen in the

flask above the water. The white cloud consists of tiny drops of water—"water-dust" it has been called—condensed by the cold air from the invisible water vapour.

In all these three experiments liquids are found to change into vapours, and then on cooling into the same liquids again. The *state* of each substance is changed, but there is no change in the *stuff* of which each consists. The solid iodine, for instance, is changed into liquid iodine, and this into the vapour of iodine, and the iodine vapour back to the solid iodine again; the substance is all the time *iodine;* it is not changed into a different substance; only its state is altered. Such changes are called **physical changes.**

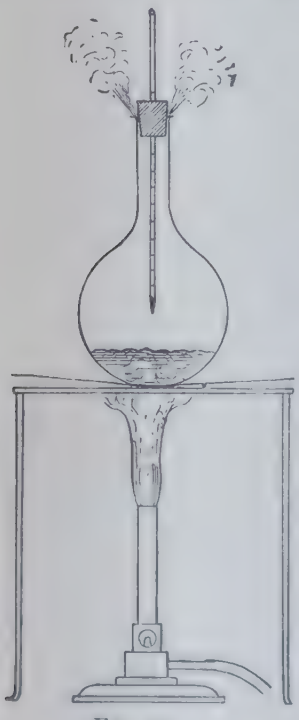

FIG. 58.

Let us study one case more fully :—

EXPT. 4. **To find the Temperature of the Vapour from Boiling Water.**—Take a large flask containing some water, and fit it with a cork. File two grooves down the side of the cork, and pierce it with a hole, through which a C. thermometer can be thrust, so that its bulb is just above the surface of the water. Place the flask on wire gauze, and heat it (Fig. 58). As soon as the water boils, and steam is issuing freely from the mouth of the flask, read the temperature six times at intervals of about one quarter of a minute. Does the temperature rise? No, *it remains at about* **100° C.**

Thus the change of state of water to the vapour of water occurs at a definite temperature, viz. 100° C. This is called the **boiling point** of water.

Is the same true of other liquids? Does each possess a definite boiling point?

EXPT. 5. **To find the Temperature of Boiling Chloroform**[1].—Take an 8 oz. flask. Support a dry test tube within it

[1] Suggested by H. P. Highton in *Practical Chemistry for Army Candidates* (Longmans), p. 124.

either by the neck of the flask, or if this is too large by a cork. Pour chloroform into the test-tube to a depth of about 2 cm. Add cold water to the flask, so that the test-tube just dips into the water. Place the flask on wire gauze, and fix a thermometer in a clamp, so that the bulb is within the test-tube, and just above the surface of the chloroform (Fig. 59). See that it does not touch the side of the test-tube. Place a small flame underneath, and drop two or three grains of sand into the chloroform, which will help it to boil evenly. As soon as the chloroform begins to boil vigorously, remove the flame, and read the thermometer five or six times. The heat of the water is sufficient to keep the chloroform boiling. What is the boiling point of chloroform? About 60° C.

Different liquids are thus found to boil at different temperatures. Moreover, when a liquid is pure, the boiling point is usually very definite and readily determined. Hence, it should be noted if possible in the case of every pure liquid.

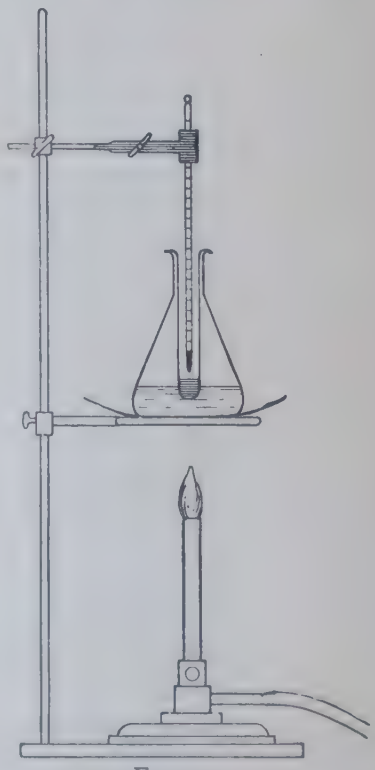

FIG. 59.

TABLE OF BOILING POINTS

Sulphur	440° C.
Mercury	357°
Toluene	110°·3
Water	100°
Benzene	80°·4 C.
Alcohol	78°·3
Chloroform	60°·2
Ether	35°

Let us repeat the vaporisation of water in another way, so as to collect the condensed water.

EXPT. 6. **The Distillation of Water.**—(1) Take an 8 oz. flask, *a*, and half fill it with water. Introduce a few pieces of broken glass or of clay pipe. Take a piece of glass tubing about 80 cm. long, and bend it about 10 cm. from one end, as shown. Fit the tube to the flask by means of a cork. Support the flask on wire gauze upon the ring of a retort-stand. Place

the other end of the glass tube within a flask, *b*, underneath a tap of cold water (Fig. 60). Heat the flask. As soon as the water has begun to boil, the steam passing down the tube is condensed on its cold surface, and when the glass tube becomes very hot, in the other flask.

The condensation of the steam, and the collection of the water produced, is spoken of as **distillation**. Taste the water. It is flat and insipid.

(2) The following is a better apparatus for the distillation of water. Take a flask, and fix in its neck by means of a cork a

FIG. 60.

thermometer and a bent glass tube, as shown in Fig. 61. Take also a "**Liebig's condenser**," which consists of a long narrow tube, the central portion of which is surrounded by a wider tube, so that cold water can be circulated between them. Select a cork which fits the tube of the condenser. Fix the cork on the tube in the flask. Place some water and a piece of clay pipe in the flask. Then, supporting the flask upon wire gauze, connect it with the condenser. Circulate cold water through the condenser. Heat the flask. When the water has begun to boil, the vapour, passing the bulb of the thermometer and down the tube, is completely condensed on the cool surface of the tube. Collect in a flask the water thus condensed. Meanwhile read the thermometer. It stands constantly at 100° C.

Non-Volatile Liquids. – Some liquids cannot be distilled, or converted into vapours which will yield the same liquids again when cooled.

EXPT. 7. **Olive Oil cannot be distilled.**—Half fill a small retort with olive oil. Fit a thermometer in a tubulure, so that the bulb is in the oil. Introduce the neck of the retort into a small flask. Heat the retort over wire gauze. Observe that the temperature rises continually higher and higher. As soon as 300° C. is reached, remove the thermometer, and replace the

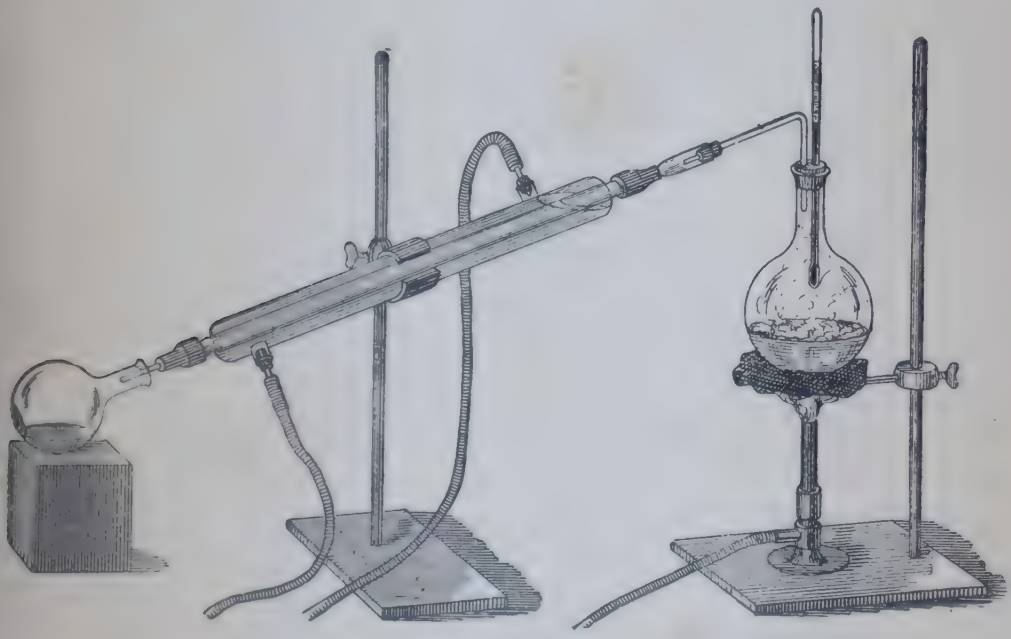

FIG. 61.

stopper of the retort. Heat further. After a time vapour is at last evolved, and being condensed in the neck, a few drops of liquid are collected in the flask. Then remove the flame. Notice the odour of the liquid in the receiver. It is very acrid, and clearly not that of unchanged olive oil. Indeed, this liquid is not olive oil at all, but another substance produced by the heat breaking up the oil.

Liquids such as olive oil, which cannot be distilled, are said to be **non-volatile**.

Exercises

1. Find the boiling points of—

(1) turpentine.
(2) benzene.
(3) pure alcohol.
(4) toluene.

2. Distil a mixture of 100 c.c. water, with 100 c.c. methylated spirit in the apparatus of Fig. 61. Read the temperature of the thermometer every minute. Collect the distillate in separate samples of about 25 c.c. each. Find the relative density of the first and last samples.

CHAPTER XI

CHANGE OF STATE. THE EVAPORATION OF WATER AND OF OTHER LIQUIDS

The Evaporation of Water.—We are often, perhaps, inclined to suppose that it is only when water is boiled that it is converted into vapour. Yet this is far from being true. Wet clothes become dry when hung in the open air, particularly on a windy day, and the puddles in the road after rain are quickly dried up in the sunshine. The quiet and invisible passage of a substance from the liquid to the gaseous state at all temperatures is called **evaporation.**

So large a part of the surface of the globe is covered with water, in sea, river, or lake, that an enormous amount of evaporation is always going on. It is evident, therefore, that the air must always contain a very large amount of aqueous vapour.

But is there any limit to the amount of aqueous vapour which the air may contain, and what conditions affect the rate at which evaporation proceeds? We must now search for answers to these questions.

Valuable information about the evaporation of liquids can be obtained by introducing them into the Torricellian vacuum of a barometer tube.

EXPT. 1. **The Evaporation of Water in a Vacuum.**—Fill two dry glass tubes, A and B, each sealed at one end, with warm dry mercury, and invert them in a trough of mercury (Fig. 62). Introduce a drop or two of water into B under the mercury by means of a pipette, C, with a curved end. The water, being lighter than the mercury, rises to the surface in

the tube. The mercury column at once falls a considerable distance. This cannot be due to the weight of the drop of water. Observe closely the top of the mercury column. The drop of water has nearly or entirely disappeared. The water has then evaporated very rapidly in the vacuum. The depression of the mercury must be due to a *pressure* which the invisible aqueous vapour is exerting, in the same way that air exerts a pressure.

Supposing we introduce some more water, will the mercury sink further? We must test this by experiment. Add a few more drops of water, so that a little water rests upon the top of the mercury column. The mercury does not sink any further after this, even though two or three drops of water are added. Measure the difference in level of the mercury in the two tubes. This difference measures the pressure which the aqueous vapour is exerting.

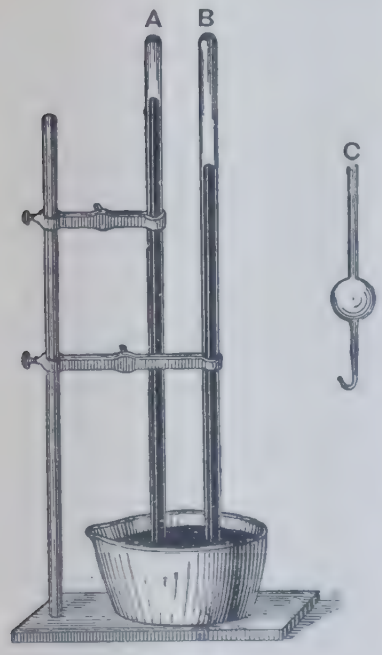

FIG. 62

From this experiment it appears that (the temperature remaining unchanged)—

(1) Water evaporates very rapidly in a vacuum.

(2) It is not possible for an unlimited amount of water to evaporate into a given space, but only a certain maximum amount. The space is then said to be **saturated.**

(3) When water is allowed to evaporate into a given space, the aqueous vapour exerts a pressure, but this pressure cannot exceed a certain maximum.

What will be the effect of an increase in temperature upon the aqueous vapour pressure? Gently warm the upper end of the tube containing the water vapour with the hand, and then with a Bunsen burner. The mercury sinks still further. Hence :—

(4) As the temperature is gradually raised, the pressure

exerted by aqueous vapour is also increased. Can we measure the pressure corresponding to any particular temperature?

Dalton's Experiment for measuring the Pressure of Water Vapour.—John Dalton was the first to carefully study this subject, and his own account of his experiments may be quoted :—

"I take a cylindrical glass tube open at both ends, and of two inches diameter and fourteen inches in length, to each end of which a cork is adapted, perforated in the middle so as to admit the barometer tube to be pushed through and to be held fast by them ; the upper cork is fixed two or three inches below the top of the tube, and is half cut away so as to admit water, &c., to pass by, its service being merely to keep the tube steady. Things being thus circumstanced, water of any temperature may be poured into the wide tube, and thus made to surround the upper part or vacuum of the barometer, and the effect of temperature in the production of vapour within can be observed from the depression of the mercurial column."

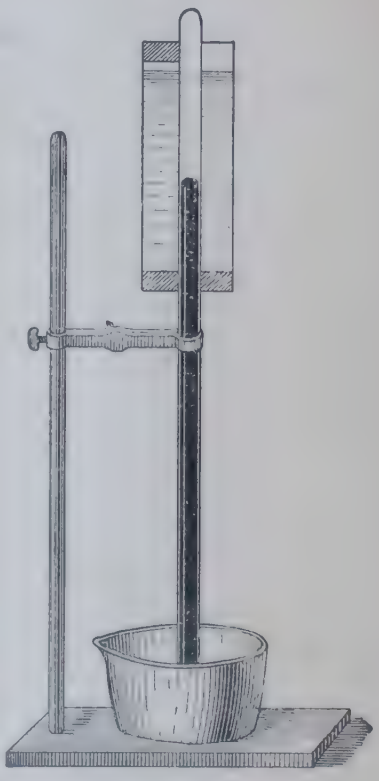

FIG. 63.

The experiments previously described show that water evaporates at ordinary temperatures below 100°. The pressure of such vapour formed at temperatures below 100° is often called **aqueous vapour pressure**, while the pressure exerted by the vapour arising from boiling water is commonly spoken of as *steam pressure*. The term *steam*, as commonly used, means either the vapour arising from boiling water, or else aqueous vapour out of contact of water heated to temperatures above 100°.

The following table shows the pressure of aqueous vapour for various temperatures, measured in millimetres of mercury, that

is to say, by the number of millimetres of mercury through which the aqueous vapour at each temperature would depress a column of mercury.

PRESSURE OF AQUEOUS VAPOUR

Temperature.	Pressure.	Temperature.	Pressure.
0°	4· + mm.	40°	54· + mm.
5	6· +	60	148· +
10	9· +	80	354· +
15	12· +	90	525· +
20	17· +	100	760· +

Saturated and Unsaturated Aqueous Vapour.—We have seen that if water be placed in a closed vessel the water evaporates very quickly if the vessel is vacuous, but that there is a certain maximum pressure corresponding to any given temperature which cannot be exceeded. If the vessel contains air, it is found that the only difference is that the evaporation is slower: gradually the maximum pressure corresponding to the temperature is attained, and cannot be exceeded. In either case the air is said to be **saturated** with aqueous vapour.

Briefly, then, at any definite temperature the air is capable of taking up only a fixed amount of aqueous vapour.

Suppose the temperature of air which is saturated with aqueous vapour be reduced, what will be the consequence? At the reduced temperature the maximum vapour pressure is less than at the original temperature. Hence some of the aqueous vapour will have to be parted with. What can become of it? Will it not be condensed as water or dew again?

We are really well acquainted with many instances of this. In a hot room there may be a large amount of aqueous vapour, yet not sufficient for it to be saturated. Close to the windows on a cold night, however, the temperature of the air will be greatly reduced, so that there will be more than sufficient aqueous vapour present for it to be saturated, and as we know a heavy dew will be deposited upon the cold glass. So also the dew often found in the morning on the grass, and on the surface of leaf and flower and cobweb, comes from the cooling of the air around them to such a temperature that the invisible aqueous vapour is in excess of that which can be contained in the air at this temperature, and it is therefore deposited.

EXPT. 2. **To show the Deposition of Dew from the Air of the Room.**—Make a freezing mixture with Glauber's salt and hydrochloric acid in a test-glass. Notice the deposition of dew on the outside of the glass.

If air containing aqueous vapour be slowly cooled, the temperature is sooner or later reached at which the air is saturated with vapour. The deposition of dew will begin at this temperature, and it is this temperature which is called the **dew-point.**

The "dew-point" is the temperature at which the aqueous vapour actually present in the air is sufficient to saturate it.

The Boiling of Water.—In the above table the number 760 mm., representing the pressure of aqueous vapour at 100° C., shows that the *pressure of aqueous vapour at* 100° *is equal to the normal atmospheric pressure.* Now it is exactly at this temperature that water evaporates in that rapid and violent way which we speak of as the *boiling* of water. This suggests that a liquid boils at a temperature such that the pressure of its vapour at that temperature is equal to the pressure of the atmosphere. How do experiment and experience bear this out?

If this be so, it follows that water should boil on a mountain at a lower temperature than at the sea level, for on the top of a mountain the atmospheric pressure is less than at sea level. This conclusion is borne out by Professor Tyndall's discovery that on the top of Mont Blanc water boils at 85°, or 15° below the temperature at which it boils at the sea level. Indeed, for every 600 feet above sea level the boiling point of water is found to fall about 1° F. But we do not need to ascend a mountain in order to alter the boiling point of water. At any given place the height of the barometer varies, and consequently the boiling-point of water is found to vary, slightly it is true, but perceptibly. It changes 1° C. for an alteration of 27 mm. in the height of the barometer.

EXPT. 3. **Franklin's Experiment to show that Water will boil below its ordinary Boiling Point when the Pressure above it is lowered.**—Fit a strong round-bottomed flask with rubber stopper, glass tube, rubber tube, and clip, as shown in Fig. 64. Introduce some water, and

support the flask over wire gauze. Open the clip, and boil the water vigorously till all the air is expelled, and its place taken by aqueous vapour. Then remove the flame, and close the clip. Allow the flask to cool until all boiling ceases. Then pour some cold water over it. The water begins to boil. The reason of this is that a large portion of the aqueous vapour in

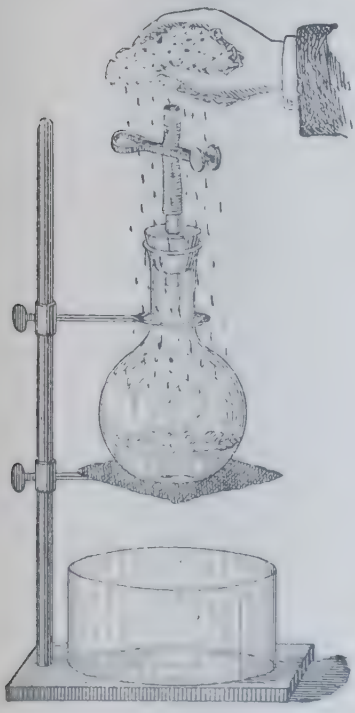

FIG. 64.

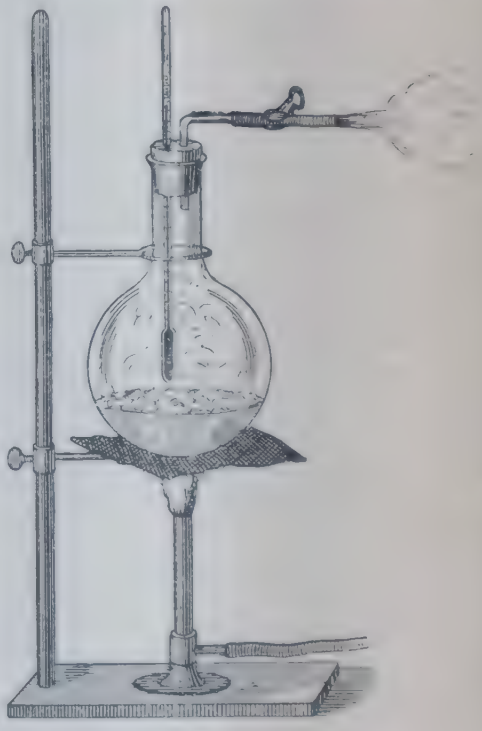

FIG. 65.

the flask is condensed through the pouring on of the cold water. Consequently the pressure on the surface of the water is so far diminished that it is less than the maximum pressure of aqueous vapour at the temperature of the water, and the water therefore **boils.**

PROBLEM

Place a flask containing hot water under the air pump, or connect it with a Bunsen pump. Can you make the water boil?

Then again, supposing the pressure above the surface of water is increased, should not the boiling point be raised?

EXPT. 4.[1] **To show that the Boiling Point of Water can be raised above 100° C.**—Pour some water into a strong flask, and close it with a rubber stopper, through which are passed a thermometer and a glass tube bent at right-angles (Fig. 65). Attach to the external end of the tube a piece of rubber tubing bearing a screw clip. Leaving the clip open, heat the flask on a flame until the water boils, and observe the temperature. Then close the clip until only a narrow exit is left for the steam. The result is that the pressure within the flask is gradually increased, and consequently the water boils at a higher temperature. *As soon as the temperature is seen to rise* 2°, release the clip at once.

Definition.—The boiling point of a liquid is that temperature at which the pressure of its vapour just overcomes the pressure of the atmosphere upon the surface of the liquid.

EVAPORATION OF OTHER LIQUIDS.

We have considered the evaporation of water in some detail. Do other liquids evaporate at temperatures below their boiling points? Can the statements made of aqueous vapour be applied to the vapours of other liquids besides water?

EXPT. 5. **To compare the Evaporation of Liquids.**—Place drops of ether, chloroform, alcohol, benzene, water, and olive oil upon a slate. Observe that some disappear or are evaporated before others.

EXERCISE

Repeat Expt. 1, using (1) alcohol, (2) ether, (3) chloroform, in place of water.

The mercury sinks further when alcohol is used than with water, and still further in the case of ether, the temperature being the same in each case. Try also in each case the influence of an increased temperature.

It is evident then that the vapour pressure of alcohol or ether is greater than that of water at the same temperature. Moreover, in every case the vapour pressure increases as the temperature rises. It follows from this that with alcohol or ether

[1] Lecture Table Experiment.

the temperature will probably be sooner reached at which the vapour pressure is equal to the pressure of the atmosphere—in other words, the temperature at which the liquid will boil. If reference is made to the table of boiling points, p. 91 it is found in agreement with this conjecture that ether boils at 35°, and alcohol at 78°·3.

LATENT HEAT OF VAPORISATION.

We have found that while the vaporisation of a liquid is proceeding by boiling, although heat is being continuously added, nevertheless there is no simultaneous rise in temperature. This heat which apparently disappears but can be made sensible again is termed the **latent heat of vaporisation**.

In the process of boiling this heat is supplied from an external source.

Is heat also rendered latent when a liquid simply evaporates below the boiling point, when no heat is supplied from a flame or fire? If heat is absorbed in such a case, it will of necessity be obtained out of the liquid itself, or from surrounding objects, and consequently their temperatures will fall. We must test this by experiment and experience.

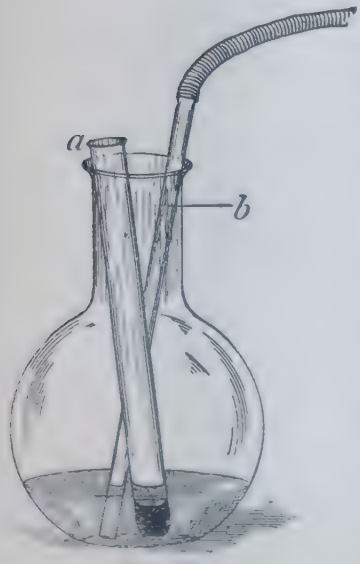

FIG. 66.

EXPT. 6. **Is Heat rendered Latent in Ordinary Evaporation?**—Pour a little ether over the hand. A sensation of cold is felt.

Again, recall the sensation of cold experienced when the body while wet after bathing is exposed to the air.

EXPT. 7.—Take a small flask, fill it one-third full with ether. Place within it a test-tube containing about 1 c.c. of water. Then, by means of a glass tube connected with an air-blast or bellows, send a rapid stream of air through the ether (Fig. 66). The ether is rapidly evaporated. Presently the water inside the tube is found frozen solid. Notice the hoar frost on the outside of the flask.

Exercise

Put some ether in a beaker. Wet the outside of the beaker with water. Stand it on a wooden block. Can you freeze the beaker to the block?

Expt. 8. **The Cryophorus.**—This is an interesting piece of apparatus devised by the physicist Wollaston (Fig. 67). It consists of a closed glass tube, bent twice at right angles, with a bulb at each end, and contains some water and water vapour, but no air. Run all the water into the bulb B. Immerse the

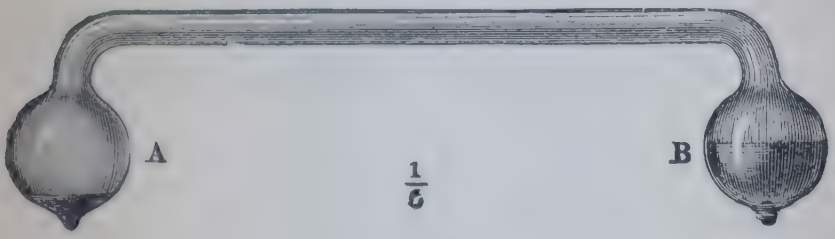

Fig. 67.

bulb A in a freezing mixture. After a time dew is seen to be deposited on the outside surface of B, and finally the water in B may even be frozen.

Expt. 9. **The Wet and Dry Bulb Thermometers.**—Take a thermometer. Wrap the bulb in cotton wool. Read its temperature. Pour a few drops of methylated spirit on the cotton wool. Wave the thermometer about in the air. Read the thermometer again. The temperature has fallen several degrees. Repeat, using (1) ether, (2) water.

It is a matter of common knowledge that the air is much more moist on some days than on others. Sometimes the air appears to be laden with moisture, and everything out of doors is dank and wet. On such a day it would be useless to hang clothes out to dry, for the air is saturated or contains as much aqueous vapour as is possible at that temperature, and if a wet thermometer bulb were moved about in such air no fall of temperature would occur. It is evident, therefore, that we can judge of the un-saturation of the atmosphere by observing whether a moist thermometer sinks below the actual temperature of the air or not. *The wet and dry bulb thermometer* (Fig. 68)

is an instrument which enables us to do this. One thermometer serves to record the actual temperature of the air. The bulb of the other is surrounded by a piece of muslin, one end of which dips into a small vessel containing water, and by these means the bulb is kept moist. On a dry day the water on the wet bulb evaporates, and a lower temperature is recorded than by the dry thermometer. Moreover, the magnitude of the difference between the readings of the two thermometers gives a measure of the degree of un-saturation of the atmosphere,

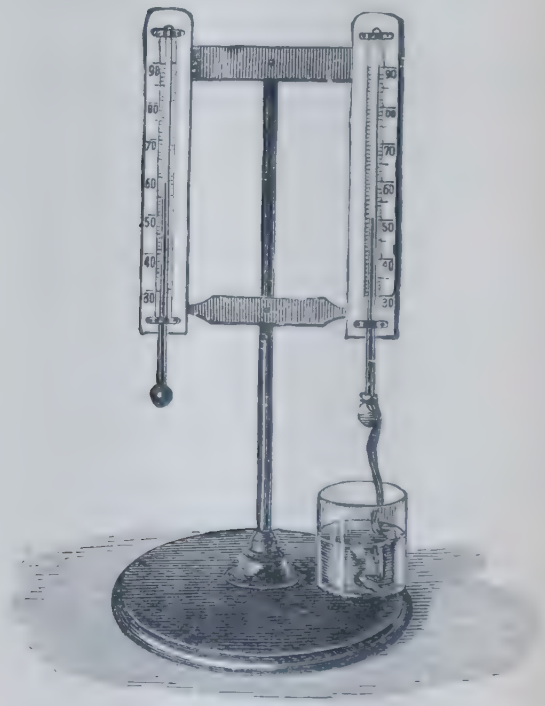

FIG. 68

for, the drier the air, the more rapid is the evaporation, and the greater the lowering of temperature. On a thoroughly damp day the two thermometers record the same temperature.

Conclusion.—It has been shown that the air always contains invisible water vapour. Sometimes the air is saturated, and the vapour is readily condensed as rain or mist or dew in the open, or as a dew on cold walls or pipes indoors. But far more often the air is not saturated, and in that case evaporation must result wherever there is water. In the open air, water is evaporated

from seas, lakes, and rivers, from the soil, and from the leaves of all trees, from the grass and all flowers, and if the heat is great, and the evaporation therefore rapid, these may fade and wither. And in a warm, closed room, which is not well ventilated, especially if it is heated by hot pipes or by a closed stove, evaporation will often take place so rapidly from growing ferns or other plants that the fronds or leaves droop and die ; evaporation of water will also take place from the surface of the skin of any occupants of the room, frequently causing exhaustion and oppression. Such a condition of a room may be prevented by not allowing the temperature of a room to exceed 60° F. by proper ventilation, and by having some water exposed in vessels in the room.

Problem

Hang up a piece of salt seaweed. Compare its condition, day by day, with the readings of the wet and dry bulb thermometer.

CHAPTER XII[1]

CALORIMETRY

The Unit of Heat.—We have already seen that when a solid is melted, or when a liquid is vaporised, heat becomes latent (pp. 85, 90 and 102), but we cannot endeavour to measure the *quantity* of heat which becomes latent until we have fixed upon some standard measure, that is upon a definite *unit* of heat.

The effects of heat are numerous, and any one of them which can be accurately measured might be used as a unit of heat; for instance, we might select as our unit of heat the quantity necessary to produce a definite increase of volume in unit mass of some standard substance. As a matter of fact, the effect of heat selected is that of increase of temperature, and the unit of heat is the **quantity of heat required to raise the temperature of 1 gram of water from 0° C. to 1° C.**; this unit is usually called **one calorie**. Hence, if 50 grams of water are to be heated from 0° to 1° C., 50 units of heat will be required.

Pending further inquiry, we shall assume that the quantities of heat necessary to raise the temperature of 1 gram of water 1° are equal wherever that degree may be on the scale of temperature; for instance, the quantity of heat necessary to raise the temperature of 1 gram of water from 80° to 81° may be taken as equal to that required to raise the temperature of 1 gram of water from 0° to 1°. It follows from this that if the

[1] *To Teachers.*—As the calculations involved in this chapter are usually felt to be rather difficult, the chapter may, if circumstances allow, be postponed with advantage to a later period in the course.

temperature of 1 gram of water be raised from 0° to 80°, 80 calories are required, and if the temperature of 20 grams of water be raised from 30° to 55°, 25 × 20 or 500 calories will be required.

It must be assumed also that, when a given weight of water, or indeed of any substance, is cooled through any range of temperature, it gives up exactly the same quantity of heat that is required to raise its temperature through the same range.

Let us next consider what occurs when we mix together two quantities of water at different temperatures.

EXPT. 1. **The Effect of mixing Two Quantities of Water at Different Temperatures.**—Support a small cylindrical vessel, *a*, made of sheet brass,[1] on corks inside a

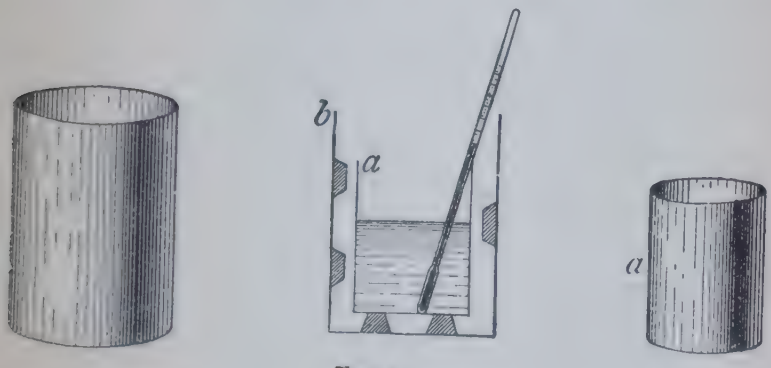

FIG. 69.

larger zinc vessel, *b* (Fig. 69). The inner vessel is called a "calorimeter" or heat-measurer. Measure into it 50 c.c. of water. Place a C. thermometer inside it. Weigh a small dry flask. Measure 50 c.c. of water into it. Place another C. thermometer inside it, and heat the flask slowly over a small flame. Meantime stir the water in the calorimeter, and read its temperature. As soon as the water in the flask is above 50°, remove the flame, stir carefully with the thermometer, quickly read its temperature, take out the thermometer, and pour the water into the calorimeter. Stir the mixture in the calorimeter, and read the highest temperature attained to one-tenth of a degree if possible. Weigh the flask, and so determine the

[1] A small glass beaker may be used if a brass vessel is not available.

quantity of water left in the flask. The following numbers were obtained in an experiment :—

Weight of warm water measured out	50 gm.
,, ,, ,, left in flask	0·5 ,,
,, ,, ,, used...	49·5 ,,
Temperature of warm water	52°·1 C.
Weight of water originally in calorimeter	50 gm.
Temperature of water in calorimeter	19°·1 C.
Temperature of mixture observed	35°·1 ,,

It is evident that the water in the calorimeter gained heat, while heat was lost by the warm water added to it. If there was no loss of heat, and no interchange of heat except between the two quantities of water, we can readily calculate what the common and final temperature should be, and this should agree with the final temperature observed, viz. 35°·1.

Probably the simplest and most convenient way is to calculate the quantities of heat which both quantities of water contained *above* that they would have possessed were their temperatures 0°. These were :—

50 × 19·1 or 955 calories in the case of the water in the calorimeter.
49·5 × 52·1 or 2578·95 calories in the case of the warm water.
Hence,
955 + 2579 or 3534 calories of heat were absorbed by 50 + 49·5 or 99·5 grams of water.

Hence, the common temperature should be $\frac{3534}{99{\cdot}5}$ or **35·5**, whereas it was actually found to be 35°·1.

The difference between the observed and calculated values is probably due to a slight loss of heat from the water; for instance, a small amount of the heat in the warm water goes to heat not only the water in the calorimeter, but also the calorimeter itself and the thermometer. (*See* Expt. 4, p. 111.)

SPECIFIC HEAT

When equal quantities of heat are communicated to equal weights of water, the temperature of each is raised by the same amount. But if equal weights of *different* substances are heated until they have acquired the same temperature, will they be

found to have absorbed the same quantities of heat? In other words, have different substances the same *capacity* for heat?

EXPT. 2. **To find whether Different Metals have the same Capacity for Heat.**—Take cylinders of lead, tin, and copper, of the same weight, say 50 grams in weight each. Suspend them by threads in boiling water for five minutes. Take three calorimeters, and measure 50 c.c. of water at the ordinary temperature into each. Place a thermometer in each. Then remove the cylinders at the same time from the boiling water, and place one in each calorimeter. Stir carefully with the thermometer, and read in each case the highest temperature attained. Have the metals heated the water to different temperatures? Which has raised the temperature of the water the most? The copper.

It follows from this experiment that if equal quantities of heat are given to equal weights of *different* substances, the temperatures of each will be raised by different amounts.

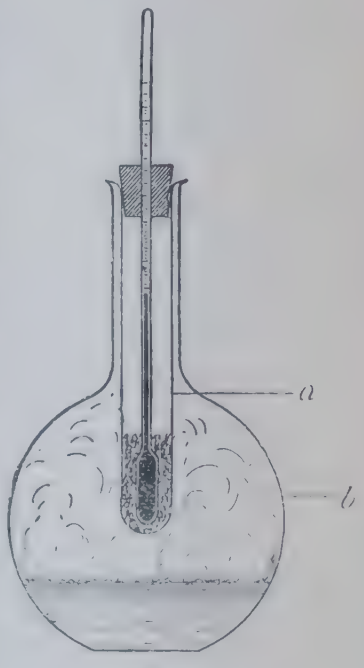

FIG. 70.

Definition of Specific Heat.—*The quantity of heat required to raise the temperature of* 1 *gram of a substance from* 0° *C. to* 1° *C. is called its specific heat at* 0°. Since, however, very nearly the same quantity of heat is required to raise the temperature of any substance 1° wherever that degree may be on the scale of temperature, we may state as approximately true that the **specific heat of any substance is the quantity of heat necessary to raise the temperature of 1 gram of it 1° C.**

EXPT. 3. **To find the Specific Heat of a Metal.**[1]—Measure into a thin brass vessel 100 c.c.[2] of water, and place a thermometer in it. Weigh into a hard test-tube, *a* (Fig. 70), 50 grams[2] of the metal. Close it with a stopper, through which is

[1] *E.g.* granular lead, iron filings, granulated zinc, copper turnings.
[2] Half quantities may be used.

passed another thermometer, the bulb of which must be surrounded by the metal. Heat the metal as nearly as possible to 100° C. by suspending the test-tube for some time in a flask, *b*, containing boiling water. Read its temperature, read also the temperature of the water in the calorimeter, and then transfer the metal as quickly as possible into the water in the calorimeter, stir with the thermometer, and observe the highest temperature reached. Thus—

Weight of metal	50 gm.
,, water	100 gm.
Temperature of metal before experiment	100° C.
,, ,, cold water before experiment	12° C.
,, ,, water after experiment	16° C.

Assume *calories lost by metal = calories gained by water;* then,

Metal fell in temperature	100 – 16 or 84 degrees.
Water increased in temperature ...	16 – 12 or 4 degrees.
Water gained...	100 × 4 or 400 calories.
i.e. metal lost	400 calories.

If 50 gm. metal cooling 84° lost 400 calories,

1 gm. ,, ,, 1° lost $\frac{400}{50} \times \frac{1}{84}$ or $\frac{8}{84}$ calories.

Hence,

Specific Heat of metal = **0·095**.

The student should observe that (1) the metal must be thoroughly heated, (2) as the difference of temperature to be measured is very small, temperatures should be read to the nearest tenth of a degree.[1]

Water Equivalent of a Calorimeter.– If warm water is added to a calorimeter, containing a thermometer, which is at the temperature of the air, the calorimeter and thermometer are heated, and on the other hand there is a loss of heat from the water, which must be allowed for in accurate determinations of specific heat. Now there must be a certain quantity of water which will absorb just as much heat as the

[1] *Note for Teacher.* – The method described above is only moderately accurate, but it is instructive, and does not take too long. The student realises at least what is meant by "specific heat," which is the important thing. It may be noted that the "water equivalent" is neglected, and that it is assumed that 100 c.c. = 100 grams.

calorimeter and thermometer absorb in being heated through any range of temperature. This equivalent amount of water is called the *water equivalent of the calorimeter*, and must be allowed for in accurate determinations of specific heat.

EXPT. 4. **To show the Absorption of Heat by a Calorimeter.**—Support a calorimeter as before on cork in an outer zinc vessel. Heat 100 c.c. water in a beaker containing a thermometer to about 35° C. As soon as this temperature is reached, remove the flame, stir well, read the temperature, at once pour the water into the calorimeter, put the thermometer in the calorimeter, and stir well, watching closely the thermometer. The temperature falls rapidly through a small range owing to the absorption of heat by the calorimeter. Read the temperature. From this point the temperature continues to fall slowly, owing to the cooling of the apparatus.

State thus :—

Temperature of water in beaker ... ... = 35°·4 C.
,, ,, ,, calorimeter ... ... = 34°·5 C.
∴ Fall in temperature ... ... = 0°·9 C.

For a fuller treatment of this subject the student should refer to books on Physics.

TABLE OF SPECIFIC HEATS

SOLIDS

Magnesium... ...	0·250	Zinc...	0·096
Aluminium	0·214	Copper	0·095
Glass	0·188	Silver	0·057
Sulphur	0·177	Tin...	0·056
Diamond	0·147	Gold	0·032
Iron	0·114	Lead	0·031

LIQUIDS

Water...	1·000	Oil of turpentine ...	0·462
Alcohol	0·615	Mercury	0·033
Ether...	0·503		

It is seen at once from these numbers that the specific heat of water very greatly exceeds that of most other substances.

See p. 117 for Examples on Specific Heat.

LATENT HEAT OF LIQUEFACTION

We have already found that heat becomes latent when any solid substance is melted, and we are now in a position to measure the latent heat of liquefaction of ice. But first we must define exactly what it is we wish to measure.

Definition of Latent Heat of Liquefaction.—*The latent heat of liquefaction of any substance is the number of calories of heat absorbed by* 1 *gram of it in passing from the solid to the liquid state without rise of temperature.*

N.B.—This is equal to the number of calories of heat evolved when a gram of the liquid substance is reconverted to the solid state without fall of temperature.

EXPT. 5. **To measure the Latent Heat of Liquefaction of Ice.**—Weigh a small sheet-brass calorimeter. Add 100 c.c. of water, and weigh again. Heat it carefully to about 20° C. Then weigh out roughly on filter-paper about 10 grams of ice. Place the calorimeter on cork supports within a zinc can, and carefully read the temperature of the water in it. Dry the ice with blotting paper. Quickly introduce it into the calorimeter, stir thoroughly, and observe the lowest temperature attained. Finally, remove the thermometer, allowing the water to drain off into the calorimeter, and weigh again. The increase in weight gives the weight of ice added. In an experiment the following numbers were obtained :—

Weight of calorimeter	21·4 gm.
,, calorimeter and water...	119·2 ,,
,, calorimeter, water, and ice...	128·5 ,,
Initial temperature of water	21°·0 C.
Final ,, ,, ...	12°·5 C.

Thus, 119·2 − 21·4 or 97·8 grams of water were cooled from 21° to 12°·5.

128·5 − 119·2 or 9·3 grams of ice were added.

This ice *was first melted, and secondly heated to* 12°·5.

Assume

heat gained by ice = heat lost by water.

Then,

heat lost by water = 97·8 × (21 − 12·5)

= 831·3 calories.

Therefore,

heat gained by 9·3 grams ice = 831·3 calories.

But the calories gained by 9·3 grams of melted ice at 0°, due to 12°·5 rise of temperature... } = 9·3 × 12·5
= 116·25

Hence, the number of calories absorbed in melting 9·3 grams of ice ... ... } = 831·3 − 116·25
= 715·05

Therefore,

number of calories absorbed in melting 1 gram of ice ... ... } = $\frac{715·05}{9·3}$ or **76·9**

The value obtained here for the latent heat of liquefaction of ice must be slightly inaccurate, owing to a small loss of heat from the apparatus, and to the ice being slightly wet on the surface. The true value lies between 79 and 80, but in round numbers it may conveniently be taken as **80**. We have arrived then at the result that in the change of 1 gram of ice at 0° into 1 gram of water at 0° no less than 80 calories of heat are rendered latent ; that is to say, as much heat as could heat 80 grams of water from 0° to 1°, or heat 1 gram of water from 0° to 80°.

It is interesting to observe the values of the latent heats of liquefaction of other bodies.

TABLE OF LATENT HEATS OF LIQUEFACTION

Ice	80	Tin	14+
Sodium nitrate... ...	63	Iodine	11+
Nitre...	47+	Sulphur	9+
Zinc	28+	Lead	5+
Silver	21+	Phosphorus ...	5+

It is very noticeable that the latent heat of ice is very great as compared with those of other substances, and this is a matter of the greatest importance. If snow and ice were immediately converted into water at 0° C. with little or no absorption of heat, the sudden melting of the snow on the hills and mountains would expose the inhabitants of the valleys below them to the most terrible floods and inundations.

LATENT HEAT OF VAPORISATION

Just as heat disappears whenever a solid is liquefied, so also we have already seen that when a liquid is vaporised heat becomes latent.

Definition of Latent Heat of Vaporisation.—*The latent heat of vaporisation of any liquid at a given temperature is the number of calories of heat absorbed by* 1 *gram of it in*

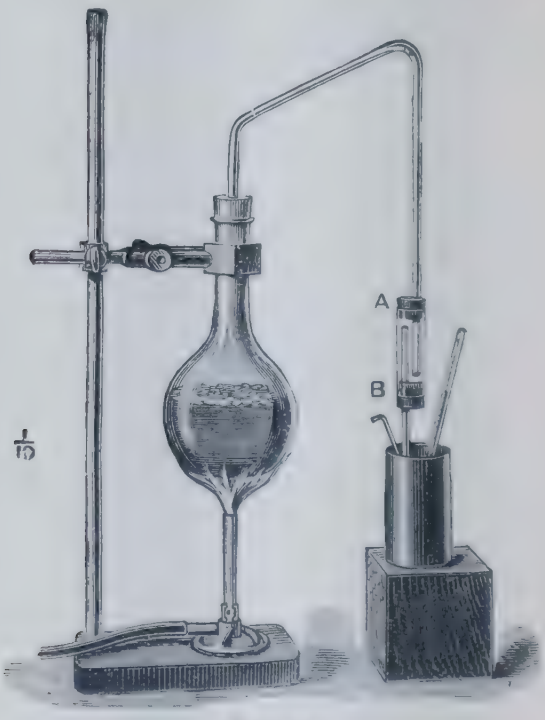

Fig. 71.

passing from the liquid to the gaseous state without rise of temperature.

When a vapour is condensed again into a liquid, the heat which was insensible or latent becomes sensible again, and the number of calories evolved is equal to the number previously absorbed.

Expt. 6. To measure the Latent Heat of Vaporisation of Water.—(1) Place some water in a 500 c.c. flask. Close it with a stopper, through which are passed a thistle-funnel (its lower end dipping below the surface of the water, not shown in Fig. 71) and a glass tube, bent as shown. Proceed to

heat the water. Also take a wide glass tube, A B, closed by stoppers, through the lower of which a straight glass tube is passed. Connect this apparatus, which is to serve as a trap, with the flask by means of the bent glass tube. The trap serves to catch the condensed water which is carried over by the steam. When the steam issues freely from the end of the tube, measure 50 c.c. of ice-cold water into a weighed brass calorimeter or beaker, and place it so that the end of the tube reaches nearly to the bottom of the calorimeter.

The steam is condensed with a loud noise by the cold water, and the temperature rapidly rises, till finally the water in the beaker itself is raised to near the boiling point. Then remove the flame, and also the beaker, and observe that the volume of the water in the beaker is only increased by about one-fifth. If the increase in weight of the beaker is 10 grams, this shows that 10 grams of steam at 100°, in changing into 10 grams of water at 100°, set free sufficient heat to raise the temperature of 50 grams of water from 0° to 100° C. What value does this result give for the latent heat of vaporisation of water?

This experiment shows that the latent heat of vaporisation of water has a very high value, and that therefore steam should never be needlessly wasted.

(2) Try to measure accurately the latent heat of vaporisation of water. For this it will be necessary to stop the condensation of the steam, when the temperature of the water reaches about 50°, because at higher temperatures its loss of heat by cooling becomes much greater. Proceed thus:—Measure 100 c.c. of cold water into a calorimeter, and weigh. Read its temperature. Pass steam in until the temperature rises just above 50° C. Then lift the delivery tube out from the calorimeter, stir well, and at once read the temperature. Thus:

Weight of cold water	= 100 gm.
Weight of calorimeter when cold	= 141·7 ,,
,, ,, ,, warm	= 148·7 ,,
Temperature of cold water	= 13°·0 C.
,, ,, warm water	= 53°·3 C.

Hence, 100 grams of water were warmed from 13° to 53°·3 C., and 7 grams of steam were condensed.

Now, the steam *was first condensed, and, secondly, cooled to* 53°·3 C.

Assume

heat gained by water = heat lost by steam.

Now,

heat gained by water = 100 × (53·3 − 13)
= 4030 calories.

∴ heat lost by 7 gm. of steam = 4030 calories.

But the calories lost by 7 grams of *condensed* steam at 100° due to (100° − 53°·3) or 46°·7 fall of temperature = 7 × 46·7.
= 326·9.

∴ the calories given out by the condensation of 7 grams steam into 7 grams water at 100° ... ... = 4030 − 326·9
= 3703·1

∴ the number of calories given out by the condensation of 1 gram steam ... ... = $\frac{3703 \cdot 1}{7}$
= **529**

The true value of the latent heat of vaporisation of water at 100° is found to be **537**; that is to say, in converting 1 gram of water at 100° into steam at 100°, no less than 537 units of heat are rendered latent.

TABLE OF LATENT HEATS OF VAPORISATION

Water	537	Oil of turpentine	74
Alcohol	202+	Bromine	45+
Ether	105+	Iodine	24
Acetic acid	102		

It is seen from this table that the latent heat of vaporisation of water is exceptionally high, and greatly exceeds that of almost every other liquid, just as the value of its specific heat and that of its latent heat of liquefaction are higher than those of other liquids.

In view of the fact that water exists in vast quantities in the liquid form on the surface of the globe, that its vapour is always present in the air, and that ice and snow are forms in which it is often found, there can be no doubt that the above exceptional properties of water are of the highest importance in maintaining a fairly uniform temperature on the surface of the earth.

Examples XII

(*a*) Specific Heat

1. Calculate in each of the following cases what the common temperature should be, assuming no loss of heat :—

i. 50 gm. water at 51° C., added to 50 gm. water at 16° C.
ii. 49·5 ,, 50° C. ,, ,, ,, 15° C.
iii. 50 ,, 95° F. ,, ,, ,, 16° C.
iv. 1 litre ,, 99° C. ,, 900 c.c. ,, 10° C.

2. How many calories of heat are required to heat—

i. 50 gm. copper from 0° to 100° C. (S. H. copper = 0·095)
ii. 50 ,, ,, 10° to 95° C. ,, ,, ,,
iii. 25 gm. lead from 20° to 100° C. (S. H. lead = 0·03)
iv. 1 kilogram iron from 20° to 100° C. (S. H. iron = 0·11)

3. Find the specific heat of a silver watch chain, having given

Weight of chain	90 gm.
Temperature of hot chain	99° C.
Volume of cold water	100 c.c.
Temperature of cold water	15° C.
Temperature of mixture	19°·1 C.

4. If 200 grams of mercury at 100° C. are mixed with 100 grams of water at 7° C., and the resulting temperature is 13° C., find the specific heat of mercury.

5. 10 grams of a certain substance are heated to 100° C. and placed in 75 grams of water at 4° C. The final temperature of the water is 10° C. Find the specific heat of the substance.

6. Find the final temperature of a mixture of 100 grams alcohol at 10° C. and 100 c.c. water at 0° C., assuming that the S.H. of alcohol is ·6, and that the only heat change is an interchange between the alcohol and the water.

7. Find the final temperature of 100 grams of water originally at 0° C., after dropping into it 10 grams of copper nails at 100° C. (S.H. of copper = ·095.)

8. Find the resulting temperature when 10 grams of water at 100° C. are poured into a copper vessel at 0° C. weighing 100 grams. (S. H. of copper = ·095.)

9. What would be the rise of temperature on dropping 100 grams of lead shot at 100° C. into 100 grams of water at 0° C.? (S. H. of lead = ·03.)

10. How many grams of silver at 90° if dipped into 100 grams alcohol at 10° would raise its temperature to 20° C.? (S. H. of silver = ·057; alcohol = ·615.)

11. What weight of mercury at 100° if mixed with 100 grams of water at 0° would raise its temperature to 10° C? (S. H. of mercury = $\frac{1}{30}$.)

(*b*) Latent Heat

12. How much heat would be required to melt 1 kilogram of ice without rise of temperature?

13. Calculate the latent heat of liquefaction of ice from the following data:—

	(i.)	(ii.)	(iii.)
Weight of water	100 gm.	70·5 gm.	100 gm.
,, ,, + melted ice	104·5 ,,	76·3 ,,	105·7 ,,
Initial temperature of water ...	21°·4 C.	21°·5 C.	21°·4 C.
Final ,, ,, ...	17°·1 C.	13°·9 C.	15°·9 C.

14. Calculate the common temperature reached on mixing 10 grams of ice at 0° C. with 100 c.c. of water at 80° C.

15. How many grams of steam would be required to raise the temperature of 1 litre of water from 0° to 100° C. (L. H. of vaporisation of water = 537.)

16. Calculate the latent heat of vaporisation of water from the following data:—

	(i.)	(ii.)
Weight of cold water	100 gm.	100 gm.
Weight of calorimeter when cold	117·3 ,,	137·7 ,,
,, ,, ,, warm	124·0 ,,	144·1 ,,
Temperature of cold water	12°·2 C.	13°·8 C.
,, warm water	50°·4 C.	50°·0 C.

17. How much heat would be required to change 1 gram of ice at 0° C. into steam at 100° C.?

CHAPTER XIII

SOLUTION

THE SOLVENT ACTION OF WATER.—What occurs when a solid substance is placed in water? Place a lump of sugar in a small beaker of warm water. The sugar crumbles down, melts, and disappears in the water.

The sugar is said to be **dissolved** in the water, or to be **soluble** in the water. And this sort of melting and disappearing in the water is called **solution**, and the water is said to have a **solvent action** upon the sugar. But the sugar is not lost, for the water tastes sweet. Is there any change in weight?

EXPT. 1. **Is there a change in weight when sugar is dissolved in water?**—Warm some water in a beaker, put in it a glass rod, and cover it with a watch-glass. Place a piece of smooth paper with a few grains of sugar on the watch-glass. Place the whole on one pan of a balance, and counterpoise it. Carefully slide the sugar into the water, and stir until it is dissolved. Is the weight altered? No.

All substances, however, do not dissolve so readily in water as sugar.

EXPT. 2. **To find whether Plaster of Paris is soluble in Water.**—Shake some powdered plaster of Paris in a small flask with water. The water becomes milky owing to powder *in suspension*. The plaster seems insoluble.

But does the evidence justify this conclusion? We know that the plaster *is not all dissolved*. Is it possible that some of it is dissolved? To settle this we must separate the liquid from the floating powder. This we can do by letting it stand, but that takes time. Can we not pour the milky liquid on some porous

material such as blotting-paper, which will allow the liquid part to run through while the floating solid particles are caught and retained on the paper? Try.

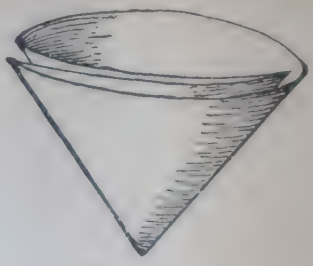

FIG. 72.

EXPT. 3. **To filter the Milky Liquid containing Plaster of Paris.**—Take a circular piece of a kind of porous paper called filter-paper. Fold it once across a diameter. Then fold it again in two. Open the folded paper so that it forms a cone, as shown in Fig. 72. Place it in a dry funnel, the cone of which is larger than the paper cone. It springs up if the paper does not fit the glass closely all round; alter the folding till it fits well. Now moisten the paper with water to keep it in place. Place the funnel in the mouth of a flask, or in the ring of a retort-stand above a clean beaker.

FIG. 73.

Take the flask containing the water which has been shaken up with the plaster of Paris. Pour the contents down a glass rod into the funnel (see Fig. 73), and take care that the liquid does not rise above the top of the paper. The plaster is retained on the paper, while a clear liquid runs through; this latter is called the **filtrate.** If the

first few drops are not quite clear, pour them on the filter again.

If the clear filtrate is evaporated, the water will be driven off. Solid plaster of Paris does not evaporate into the air. It seems probable, then, that if the clear filtrate be evaporated, solid plaster of Paris will be left behind, if the water has really dissolved any of it.

EXPT. 4. **To find whether the Clear Filtrate contains Dissolved Plaster of Paris.**—(1) Place a drop of the filtrate on platinum foil. Hold the foil by a pair of tongs over a flame.

FIG. 74.

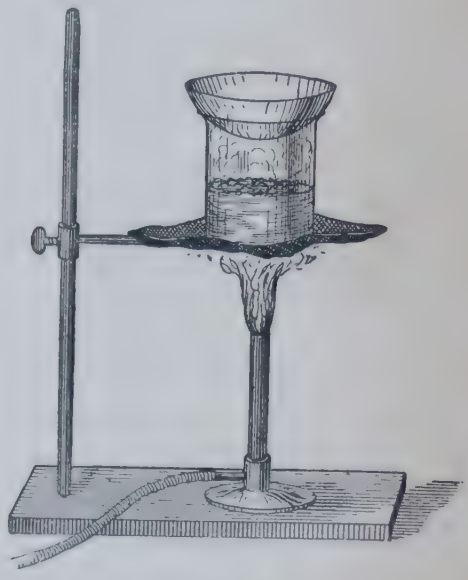

FIG. 75.

A white stain is left. It is clear, then, that plaster of Paris is **slightly soluble** in water.

(2) Pour some of the filtrate into a glass basin. Place the basin on a sand-bath (see Fig. 74), and heat it so that the liquid is evaporated without spurting. The volume of liquid gradually lessens, and a white deposit crusts the basin.

(3) A still surer method for avoiding loss by spurting is by evaporating the liquid on a "**water-bath.**" The basin is in this case put on a beaker in which water is being boiled, as shown in Fig. 75.

Problems

Examine the solvent action of cold water upon powdered nitre, salt, green vitriol, garden soil, gravel, chalk, Epsom salts.

SOLIDS IN NATURAL WATERS.—Do your results not bring out the fact that several substances which occur in nature are soluble in water? Then will not water which has flowed over or through the crust of the earth be found to have dissolved something? We can readily find out whether this is so.

Expt. 5. **To find whether Spring Water contains Dissolved Solids.**—Take a clean glass basin.[1] Measure 10 c.c. of clear spring water by a pipette into it. Evaporate the water by placing the basin on a sand-bath or water-bath. After the disappearance of the water, a slight solid residue is seen to be left.

Problems

Find whether rain water, river water, sea water,[2] do or do not leave a solid residue on evaporation.

You have probably found that while the rain water leaves little or no residue the sea water leaves very much more than the spring or river water.

Then *rain* water—*i.e.*, water which has not been in contact with the surface of the earth—is almost entirely free from dissolved solid impurity. *Spring* and *river* waters, on the other hand, are waters which have trickled through or over various strata of the earth's crust, so that it is probable that much of the solid matter they contain has been dissolved out of the rocks and soil over which they have passed. This will account, too, for *sea* water containing so much dissolved matter. For the rivers flow down into the sea, carrying with them that which they have dissolved out of the earth, and as the water of the sea evaporates the solid material remains behind, just as in the above experiment. A *mineral* water is a spring water which contains an exceptional amount of dissolved solids.

Observe the following table. Why do some river waters contain more dissolved solids than others?

[1] Or a watch-glass, 3″ diameter.
[2] Sea water is supplied by most chemical firms.

SOLIDS IN GRAMS PER 100,000 C.C. OF NATURAL WATERS.

Rain ... average	3 grams.	Lakes :—	
Rivers :—		Loch Katrine ...	28 grams.
Neva	5·5 ,,	Thirlmere ...	27 ,,
Irwell near source	8 ,,	Mineral Waters :—	
,, below Manchester ...	56 ,,	Vichy	503 ,,
Rhine	16–32 ,,	Eau de Seltz ...	407 ,,
Thames, at London,	40–45 ,,	Sea Water :—	
Jordan	105 ,,	Black Sea ...	1,770 ,,
Nile	158 ,,	Indian Ocean ...	3,400 ,,
Springs and Wells :—		North Atlantic Ocean... ...	3,849 ,,
Average of 200 samples	28 ,,	Dead Sea... ...	17,000 ,,

We have seen that a solid residue is left when a water containing solid matter is evaporated. If, as we have really assumed, the whole of the dissolved matter is left behind, the water which evaporates must be free from such impurity.

EXPT. 6. **Can Pure Water be obtained by Distillation?**—(1) Distil about 100 c.c. of sea water from a flask, as in Expt. 6, p. 91. See that the water does not boil violently. Throw away the first 10 or 20 c.c. of the distillate, which will have washed out the tube and flask, and may therefore contain some dust. Then collect some more, and find whether on evaporation 10 c.c. leave as much residue as before. Extremely little or none.

(2) Colour the rest of the sea water with a few drops of ink, and distil the coloured water. Is the distillate colourless? Yes.

We have, therefore, in the process of distillation a means of obtaining pure water from such water as sea water, which contains much dissolved solid matter, and is unfit for drinking purposes.

Rain Water is Distilled Water.—Distillation is always occurring in nature. From the wide expanse of the waters of the ocean, evaporation is continually taking place. Leaving the dissolved matter behind, the water vapour ascends, to descend again in the form of pure rain water upon the surface of the earth. The rain which falls towards the end of a downpour in the country far from the smoke of cities is the purest water in nature.

HARD AND SOFT WATERS.—The differences which are found between the amounts of solid matter contained in various natural waters are notable, and of much importance.

Waters which yield a lather with difficulty with soap, so that much soap is used up before its cleansing action can be utilised, are called **hard**.

EXPT. 7. **To make a Soap Solution.**[1]—Cut a piece of Castille soap so that it weighs 10 grams. Cut it into thin slices, and place them in a large flask. Add 400 c.c. methylated spirits. Warm gently, and shake until the soap is dissolved. Filter the solution into a litre flask. Wash out the other flask with a little alcohol. Add spirits so as to make the volume up to 1 litre. Store the solution in a Winchester pint bottle.

EXPT. 8.[2] **To compare the Hardness of Different Waters.**—Fill a burette with the soap solution. Take a 6-oz. glass-stoppered bottle. Measure into it 50 c.c. of the laboratory tap water. Read the burette. Run into the bottle 1 c.c. of soap solution. Close with the stopper, and shake well. Place the bottle on its side on the bench. Is there a permanent lather? No! Then, add another 1 c.c. and shake, and proceed in this way until a lather remains right across the surface of the water for three minutes. Read the volume of soap solution added.

Find in the same way what volumes of the same soap solution must be added in order to produce a permanent lather with different kinds of water—*e.g.*, (*a*) rain water, (*b*) a river water, (*c*) a spring water.

To quote some results :—

50 c.c. of rain water		required	$2\frac{1}{2}$	c.c.	soap solution.
,,	a certain spring water	,,	$3\frac{1}{2}$	,,	,,
,,	another spring water	,,	9	,,	,,
,,	a river water	,,	$11\frac{1}{2}$	,,	,,

Now the rain water had been found to contain very little dissolved solids, while the river water contained a considerable quantity. It might appear, therefore, that *hard* waters are waters that contain much solid matter. But the two spring waters had

1 Lecture Table Experiment.

2 This and other experiments may with advantage be performed by a few members of a class before their fellows.

been found to contain about equal quantities of dissolved solids, and therefore, since one of these was nearly three times as hard as the other, it would seem likely that the hardness of a water depends not only upon the quantity of dissolved solids, but also upon their nature. Test this :—

PROBLEM

Shake up rain water with a little of each of the following naturally occurring substances: take as much of the powdered substance as would cover a sixpence. Filter, and find the quantity of soap solution required to form a lather with 50 c.c. shaken with

(*a*) soda, (*b*) gypsum, (*c*) Epsom salts.

Again, some results may be quoted :—

50 c.c.	rain water	shaken with	soda	required	2 c.c.
,,	,,	,,	gypsum	,,	50+c.c.
,,	,,	,,	Epsom salts	,,	50+c.c.

It is clear, therefore, that the presence of soda in a water will not make it hard; the spring water referred to above which required only $3\frac{1}{2}$ c.c. soap solution to form a lather as a matter of fact contained soda. On the other hand, chalk, or gypsum, or Epsom salts, will render a water hard. And it is the experience of those who live in chalk or limestone districts that the well water is very hard. (There will be a further discussion of the hardness of waters containing chalk later: see Vol. II., A Research on Chalk.)

It is evident that hard waters are ill adapted for washing purposes, and that the softness or hardness of a town's supply of water will greatly affect the annual consumption of soap in every household.

EXPT. 9. **A Hard Water when boiled deposits a Solid Crust.**—Take some water which you know to be very hard. Half fill a perfectly clean 8-oz. flask with the water. Boil steadily for some minutes. Then pour out the water, and examine the inner surface of the flask. A thin film of solid matter has been deposited wherever the water was in contact with the glass.

Examine the inside of a kettle in which hard water has frequently been boiled, and note the fur on the sides.

Thus it is clear that a boiler in which immense quantities of water are heated will in the course of time become coated with a thick solid crust, if the water is hard. This leads sometimes to dangerous accidents, because the iron, if protected by the incrustation from the water, becomes greatly overheated. Hence, a water which contains a considerable quantity of solid matter in solution is generally ill suited for use in boilers.

It may be added that soft waters are much better adapted than hard waters for cooking purposes, whether for making tea and bread, or boiling vegetables and meat. On the other hand, it is found that soft waters have the disadvantage of attacking metals, especially lead, much more than hard waters do.

Problems

Find the volume of soap solution required to form a permanent lather with 50 c.c. of a hard water which has been treated in each of the following ways :—

1. Heated, and then tested while still hot.
2. Boiled for a few minutes, and then filtered and cooled.
3. Distilled.
4. Shaken with a few crystals of soda, and then filtered.

Do you not find that in each case the water has been rendered softer? The use of soda crystals in the laundry can now be understood.

Make a summary of all the measurements you have made with the soap solution.

THE SOLUBILITIES OF SOLIDS IN WATER.—Water, we have seen, can dissolve many different substances. Are all substances, however, soluble to the same extent in water? Sugar is readily soluble ; plaster of Paris only slightly so. But we must make actual measurements with different substances.

Definition of Solubility.—*The solubility of a substance in water at any given temperature is measured by the weight of the substance which can be dissolved in* 100 *grams of water at that temperature.*

EXPT. 10. **To find what Weight of Salt can be dissolved by Water.**—(1) Half fill a 6-oz. glass-stoppered bottle

with cold water. Add some common salt, and shake well. If it is entirely dissolved, add more until some remains undissolved. Note that it is impossible to dissolve an unlimited amount in the water. At the temperature of the experiment the water will only dissolve a certain amount of the salt. The solution is then said to be **saturated**. Allow it to stand. If the undissolved salt does not quickly settle, filter the solution. Weigh an evaporating basin or clock-glass. Carefully suck up 10 c.c. of the clear solution into a pipette, and transfer to the basin. Evaporate the solution to dryness on a sand-bath. Weigh. Calculate the weight of salt dissolved in 100 c.c. of the solution. In an experiment,

10 c.c. solution contained 3·04 gm. salt.

(2) Notice, however, that in the last experiment we measured the weight of salt dissolved in 10 c.c. of the *solution*, not in 10 grams of water. In order to find the "solubility," we must also *weigh* 10 c.c. of the solution in a flask. Thus,

10 c.c. solution weighed 11·80 gm.

Hence,

11·80 − 3·04 or 8·76 gm. water dissolved 3·04 gm. salt.

Therefore, 100 gm. water dissolved $\frac{3{\cdot}04}{8{\cdot}76} \times 100$ gm. salt

$= \mathbf{34{\cdot}7}$ gm. salt.

Problems

Find what weight of each of the following substances can be dissolved by 100 grams of water, saturated at the ordinary temperature :—powdered nitre, garden soil, powdered gypsum.

Do not your results bring out the important fact that substances of common occurrence differ greatly as regards their solubility in water?

The Influence of Temperature upon Solubility.—Thus far we have only examined the solubilities of substances in water at the ordinary temperature. But is it not possible that at higher temperatures the solubility of a substance will be different from what it is at a low temperature? We must, then, study the influence of temperature.

EXPT. 11. **Is Nitre more soluble in Hot than in Cold Water?**—One quarter fill a test-tube with water. Add sufficient powdered nitre to saturate it at the ordinary temperature. Then heat the test-tube in a flame. The undissolved nitre dissolves. Add more and more nitre. It can also be dissolved. Hold the test-tube in a stream of cold water from a tap, and at the same time scratch the inner surface of the tube with a glass rod. As the temperature of the solution falls, some of the dissolved nitre separates out in crystals.

Then, there can be no doubt that nitre is *more soluble in hot water than in cold water.*

PROBLEMS

Find whether powdered blue vitriol, alum, bichromate of potash, chlorate of potash, common salt, gypsum, are more soluble in cold or in hot water.

All are much more soluble in hot water than in cold water, except salt, which is scarcely more soluble in hot than in cold water, and gypsum which is even less soluble in hot water than in cold.

Solubility Curves.—The results of careful experiments on the amounts of nitre or other solids which can be dissolved in 100 grams of water at different temperatures can be best expressed by a graphical method. See Chapter XVIII, p. 179.

Solutions of known Strength.—It is sometimes necessary to prepare a solution of known strength, that is, one containing a known quantity of a substance in a given volume. The method of doing this may be illustrated by an example.

EXPT. 12. **To prepare a Solution containing 50 grams of Nitre per Litre.**—Your object will be to make the solution of the exact strength required without needless loss of time. Grind the nitre to powder. This will help it to dissolve. Heat a little gently in a dry test-tube. If any moisture condenses on the side of the glass, this shows that the nitre is damp. In this case place the nitre in a basin over a small flame. Stir with a glass rod until dry. If the nitre is weighed on the balance pan or even on a sheet of paper, some may be lost in transferring it to the vessel in which the solution is made.

Counterpoise an empty beaker. Weigh into it 5 grams of nitre. Add about 40 c.c. of hot water. You will naturally have saved time by putting water on to boil whilst weighing; besides, any attempt to boil cold water in a beaker at the bottom of which there is a thick layer of undissolved solid is liable to crack the glass. [Why?] Shake and warm until all the nitre is dissolved. Pour the solution into a 100 c.c. measuring flask. Rinse the beaker three times with warm water into the flask. The liquid is now hot, and will contract on cooling. Run water over the outside of the flask to cool it. Dilute carefully to the mark with cold water.[1]

Is there any objection to filtering a solution before it is diluted? Could you get over the objection?

If a solution has to be both cooled and diluted, which should be done first, and why?

Latent Heat of Solution.—You may have observed when some solids have been dissolved in water that the solution has appeared to become colder than the water was at first. Test this by means of a thermometer.

EXPT. 13. **To find whether the Temperature is lowered when Nitre is dissolved in Water.**—Take 30 grams of powdered nitre. Measure 100 c.c. of water into a tumbler, and read its temperature. Add the nitre, and stir rapidly with the thermometer. Read the lowest temperature reached. State thus:—

Temperature of water = 59° F.
" " solution = 45° F.
Fall in temperature = 14° F.

The tumbler feels chilled, and dew may collect on its outer surface.

How may this absorption of heat be accounted for? It will be recalled that heat becomes latent when a solid is *melted* (see page 87). And this experiment seems to show that when nitre is broken up by being *dissolved* in water, heat is also required, which can only be obtained from the liquid itself, and so its temperature is lowered. Some of the sensible heat is rendered *latent*.

[1] Ask your teacher to test the solution you have made with a "Twaddell's hydrometer," to see whether it is of the right strength.

Problems

Find whether there is any change of temperature when the following are dissolved in water. Take 30 grams of each and 100 c.c. of water: – Sal ammoniac, potassium chloride, salt, soda crystals, ammonium nitrate, Epsom salts.

DISSOLVED AIR IN WATER.– Have you not noticed the appearance of bubbles on the sides of a beaker or flask in which water is being heated, before the water has begun to boil? These look like air-bubbles, and if they are such, the air must have been dissolved in the water. And if so, the air must be less soluble in hot water than in cold water, while the reverse is almost always the case with solids. This suggests heating a large quantity of water till it boils, in order to see whether any air is expelled.

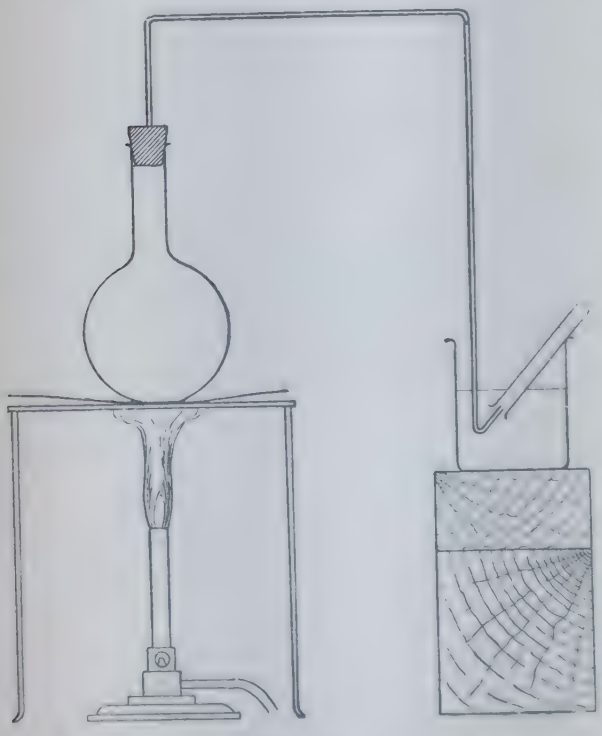

Fig. 76.

Expt. 14. **To find whether Air can be expelled from Water by Heat.**—Fill a 16-oz. flask completely with water (preferably from a spring. Why?) Fit it with a rubber stopper, and a glass tube bent as in Fig. 76, one end of which is flush with the lower surface of the stopper. Fill the tube with water, and then push the stopper well into the neck of the flask. Support the flask on wire gauze. Let the other end of the tube dip into water in a tumbler. Fill a test tube with water and invert it above the mouth of the tube (Fig. 76). Proceed to heat the flask until the water

boils. Air is carried over into the tube. After boiling a few minutes, remove the burner.

Measure the volume of air collected when cool, and also the volume of the flask. Calculate the volume of air obtained from 100 c.c. water. A certain spring water was found to contain 3 c.c. air per 100 c.c. water. What do you find?

Taste some of the water which has been boiled in the flask. It is very flat and insipid. It seems likely that this is *due to the loss of the dissolved air*. Now drinking water is often obtained at sea from sea water by distillation; but such condensed water is very unpalatable, unless it is afterwards exposed in thin streams to the air. If such water be then heated as in the last experiment, air is given off again.

Moreover, since atmospheric air is necessary for the life of man, it seems probable that it is the *air dissolved in water which supports the life of fishes and of all aquatic animals*. This is confirmed by the fact that no fish can live in water which has been boiled.

We may, therefore, conclude that unless water is boiled or distilled it will contain dissolved air, and that the small amount of air which is found dissolved in water is really obtained from the atmosphere.

SOLVENT ACTION OF OTHER LIQUIDS.—We have up to the present studied the solvent action of water alone upon solids. But though water is by far the most important and common of liquids, there are many others in frequent use; for example, ether and methylated spirits. Will the same substances dissolve in them as in water? Are those substances which are insoluble in water also insoluble in these liquids? Try.

EXPT. 15. **To try the Solvent Action of Ether on Tallow.**—We know from daily experience that fats and oils do not dissolve in water. But will tallow, for instance, dissolve in ether? Fill a burette with ether.[1]

Place a small lump of tallow in a dry test-tube. Add about 2 c.c. of ether. Shake, and stir with a glass rod. The tallow is dissolved.

[1] *Caution*: Take care that no flame is near, since ether vapour is very inflammable. Members of a class may conveniently obtain ether from a burette.

The experiment shows that the solvent action of ether is very different from that of water. It is probable, therefore, that other liquids will also be found which will dissolve substances though they may be insoluble in water. Carry out Exercises 9 and 10 below. Tabulate your results.

Problems and Exercises

Solution

1. Find whether the substances A, B, C, D[1] are

   (*a*) soluble in cold water;
   (*b*) more soluble in hot than in cold water.

2. Find whether the following are soluble in water. In each case state the evidence on which you form your conclusion:—

Sulphur, bleaching powder, pyrogallic acid, permanganate of potash, quick lime.

3. Prepare solutions in water containing 50 grams per 1000 c.c. of

Common salt; Bichromate of potash.

4. Measure out 50 c.c. of each of the solutions made in Exercise 3. into a glass basin, and find the weights of the dissolved substances.

5. Prepare a solution of common salt containing 58·5 grams per litre.

6. Prepare a solution of silver nitrate containing 17 grams per litre.

7. From the solution made in Exercise 5 prepare a weaker solution containing 5·85 grams per litre.

8. From the solution made in Exercise 7 make another solution containing 5 grams per litre.

9. Find whether the following will dissolve in ether at the ordinary temperature:—

Paraffin Oil	Iodine.
Rosin.	Water.
Olive Oil.	Methylated spirits.
Lard.	Nitre.
Camphor.	Salt.

10. Find the solvent action of (1) methylated spirits, (2) benzene on the substances given in Exercise 9.

[1] *E.g.* $Pb(NO_3)_2$ KCl, Rochelle Salt, silica.

CHAPTER XIV

CRYSTALLISATION

CRYSTALLISATION.—In by far the larger number of cases the solubility of a substance in water is increased by increase of temperature. Consequently, in general, if hot water be saturated with a substance, some of this substance will separate out in the solid state when the solution is cooled. In many cases it will separate out in crystals.

EXPT. 1. **To recrystallise Nitre.**—Measure about 100 c.c. of water into a small flask and heat it. Weigh out about 60 grams of nitre. Powder the nitre and dissolve it in the hot water. Filter the hot solution into a crystallising dish, and allow it to cool gradually. When cold, pour off the solution from the crystals which have formed. Transfer the crystals to some filter-paper, and dry them by gently pressing them with other pieces of clean paper. What weight of crystals have you obtained? About 40 grams should be obtained. Select a good large crystal or bundle of crystals, and draw it twice the actual size.

Circumstances affecting Crystallisation.—Perfect crystals are hard to obtain. The conditions of their growth are shown by the following experiments :—

EXPT. 2.—Prepare a warm saturated solution of nitre in a test-tube. Divide it into two equal portions. Put one aside to cool slowly. Cool the other rapidly under the tap. Shake it, and rub the inside of the test-tube with a glass rod. A fine crystalline precipitate falls. In the first portion larger, pointed, needle-like crystals will slowly form. In fact, when a solution cools slowly, large crystals may form; shaking and rapid cooling tend to produce small crystals.

EXPT. 3.—Dissolve some crystals of Glauber's salt (sodium sulphate) in a very little hot water. Filter into a clean flask. Close the mouth of the flask with a plug of cotton-wool to exclude dust. Allow the solution to become quite cold, and avoid shaking it. It does not crystallise. Remove the plug Let fall some tiny specks of solid Glauber's salt into the solution. Feathery crystals start from where the dust fell, and in a few seconds the flask contains a solid mass of crystals. This shows that dust particles tend to act as centres round which crystallisation may start, and that from a strong solution crystals form all massed together.

EXPT. 4. **To prepare good large Crystals of Alum.**—Prepare a strong hot solution of alum by dissolving 15 grams of alum in 75 c.c. of hot water. Cool rapidly to make sure that it is strong enough to deposit crystals. (1) If a great many crystals are formed, pour off the liquid with only a few of the crystals. Heat gently until the crystals are dissolved. Filter the liquid into a crystallising dish. You have now secured a solution which will be *saturated whilst warm*, and able to deposit a *small* quantity of crystals on further cooling. Place a piece of paper over the dish, and put it away in a cool quiet place. Examine in an hour's time. If many crystals have formed, the solution is still too strong. Suppose there are a few crystals. Leave them for twenty-four hours. Then pick out the six best. Transfer them with a glass rod to another dish. Pour the cold "*mother-liquor*" from which they were deposited gently over them, and put aside in a cool place. Arrange them so that they are not in contact. Try also suspending a crystal by a thread looped around it in some of the mother-liquor. After twenty-four hours examine them again. Remove any tiny crystals adhering to them. Turn them over so that they may rest on different faces. If the alum solution is able to evaporate slowly, it will go on enlarging these crystals. Notice that they are of an entirely different form from that of the crystals of nitre.

(2) If the solution is too weak to deposit crystals, it must be concentrated, either by heating it, or by leaving it to evaporate for several days into the air. If, by heating it, too much water is boiled away, the crystals which form on cooling may be all massed together. To find out when the solution has been

boiled down just enough, take out a single drop of it on a *cold* glass rod, and wait to see whether this drop crystallises. Do not let the rod get hot in the solution, for if so the drop removed would evaporate on the hot rod, and crystallise before the bulk of the solution was ready to do so.

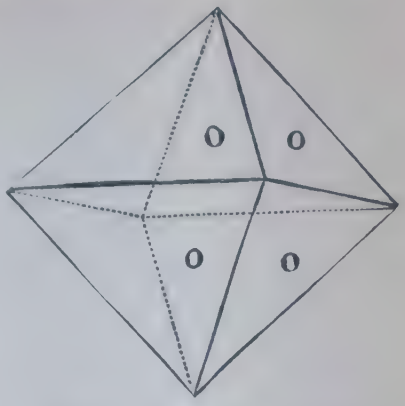

FIG. 77.—PERFECT CRYSTAL.

The shape of a perfect alum crystal is that of the regular solid called the *octahedron*, bounded by eight flat sides or *faces*, each an equilateral triangle (see Fig. 77). In less perfect specimens two of these faces are often enlarged at the expense of the others.

Almost every chemical substance has a certain shape or shapes in which it can crystallise. Draw good-sized crystals

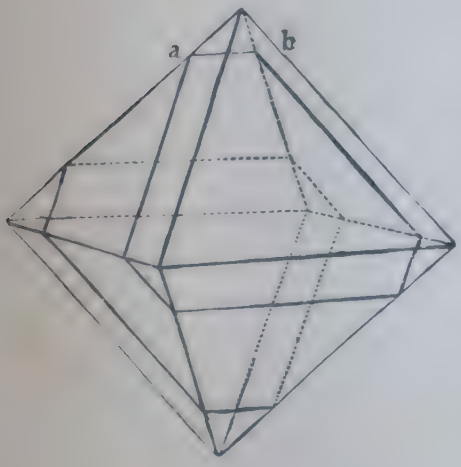

FIG. 78.

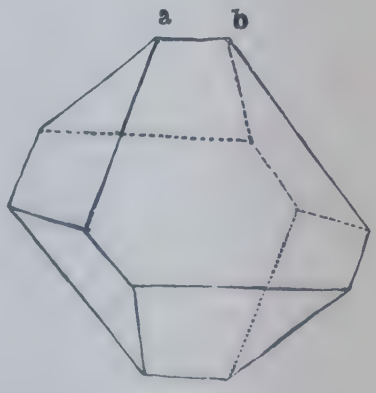

FIG. 79

of rock-salt and of calcspar. The crystals of salt are *cubes*, those of calcspar *rhombs*. On the other hand, crystals of nitre are portions of six-sided *prisms*, with bluntly-pointed ends.

Exercises

1. Cut a soap model of a sugar crystal.
2. Cut a cube of soap. Mark the centres of the faces with pins. Cut away the corners to the pins so as to leave an octahedron.
3. Cut parallel slices off two opposite faces of the octahedron of soap, as indicated in Figs. 78 and 79.
4. Make a cardboard model of an octahedron.

WHAT IS A CRYSTAL?—But what is the distinguishing feature of a crystal? It is a solid of definite shape bounded by flat sides or faces which reflect the light. But does a crystal differ from a piece of cleverly cut glass? The following experiment may suggest an answer.

Expt. 5. **To show a Difference between a Crystal of Rock-Salt and a Piece of Glass.**—Take a good crystal of rock-salt. Break it roughly with a hammer or in a strong mortar. Examine the fragments. They are all either little cubes or portions of cubes. Now break down a piece of glass rod in the same way. The fragments are of all shapes, and they appear to have no relation to one another.

Exercises

1. Crush a crystal of calcspar, and describe in words the shape of the fragments.
2. Find whether crystals of potassium ferrocyanide are equally easily cut in all directions.

The crystal of rock-salt seems then to possess definite lines along which fracture takes place, while the glass is broken as readily in one direction as in another. These lines along which the crystal can be readily broken or cut are called *lines of cleavage*. A characteristic of a crystal seems to be that it possesses a **definite structure**.

The existence of this definite structure is also impressed upon us by the fact that all the crystals of a given substance belonging to the same general form, for example crystals of calcspar, which are rhombs, have the *corresponding angles equal*.

EXPT. 6.—**To show that the Corresponding Angles of Crystals of Calcspar are Equal.**—Cut two slips of cardboard. Fasten two ends together by piercing a pin through them, so that they can turn about the pin as a pivot. Choose two good crystals of calcspar. Place them on the bench in the same relative position. Take up one of the crystals, and place it in the angle between the two slips of cardboard, so that each slip lies along a face of the crystal (Fig. 80). Hold the slips in this position. Now take up the other crystal, and see whether the *corresponding* angle in it will fit exactly into the angle between the slips of cardboard. It is found to do so.

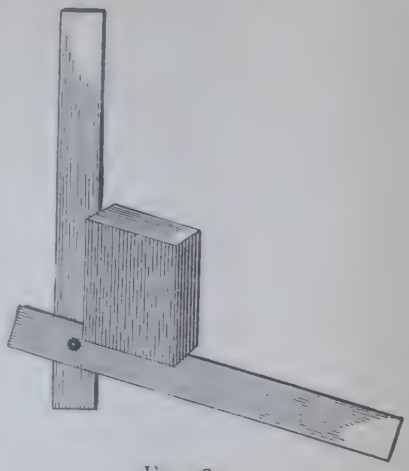

FIG. 80.

EXERCISES

1. Compare other corresponding angles on the same two crystals of calcspar.

2. Compare a third crystal of calcspar with one of those already taken.

3. Compare corresponding angles of crystals of

(*a*) rock-salt.
(*b*) alum.
(*c*) chrome-alum.

It is found that however the general shape of a crystal may be distorted *the corresponding angles remain constant.*

This is well shown in snow-crystals, which assume many fantastic shapes, but they are all built up from a simple star whose rays intersect at 60°.

The thought will readily occur that it is the *structure*, spoken of above, which determines the outward form. And it might be supposed that a cube of rock-salt consisted of a very large number of very small cubes; but this is not necessarily the case.

EXPT. 7. **To build Geometrical Shapes with a Number of Spherical Balls.**—(1) Take a number of fives-balls or marbles, and pack them in a square so that one ball is touched by four others. To prevent the balls rolling, enclose a square by a cardboard rim fastened at the corners. Pile others in the hollows between the balls. Then add yet another layer of balls, and so on. Finally, half a regular *octahedron* is the result. Cannon balls are sometimes piled like this.

(2) Arrange the balls in a triangle, the innermost ball being touched by six others. Place balls in the hollows as before. The result is the corner of a *cube*.

A cubical crystal is, therefore, not necessarily built up with cubes.

On the other hand, it may be stated that there is every reason to believe that the forms of crystals *are* due to the symmetrical arrangement of the tiny particles which compose them.

SEPARATION AND PURIFICATION OF SUBSTANCES BY MEANS OF CRYSTALLISATION.—We have found that substances differ greatly in solubility in water. It follows, therefore, that many mixtures can be entirely separated into their constituents by treatment with water. This will be most complete if one portion of a mixture is soluble in water while the rest is quite insoluble.

EXPT. 8. **To separate a Mixture of Sand and Nitre.**—*i.e.* to obtain from it all the sand dry and free from nitre, and some of the nitre dry and free from sand.

Grind some sand and nitre together in a mortar. Place about 10 grams of the mixture in a beaker. Add some *hot* water. Boil. The nitre dissolves, while the sand remains undissolved. The finer particles of sand are floating in the liquid, but would in time settle. We could then pour off or *decant* into a second beaker a liquid which would be a solution of nitre, muddy with the finest sand, leaving in the first beaker sand wet with a solution of nitre. This would be a rough but not a complete separation. The sand must be washed, and the solution of nitre filtered.

Filter the muddy liquid. It is probably still hot, which is an advantage, since a liquid filters most quickly while hot.

(1) *To recover the Nitre from the Solution.*—If we only require *some* of the nitre, we need not wait for it all to run through the filter. Evaporate the clear nitre solution in a porcelain basin on wire gauze until it is sufficiently concentrated to deposit crystals; then pour the solution into a small beaker, and cool it by floating it in a tumbler containing cold water. Dry the crystals on filter-paper.

(2) *To clean the sand*, which was left wet with the nitre solution. Pour a little hot water on the sand in the beaker. Shake. If we wish to get all the sand, this washing must be run through the filter-paper. Wash similarly several times, transferring the sand to the paper. Some of the sand clings obstinately to the side of the beaker. Use the jet of the wash-bottle to dislodge the sand, and wash it completely on to the filter. Wait till the liquid has completely run through the paper. Gently remove the filter-paper from the funnel. Open it out flat, and place it on a pile of four filter-papers. Place the pile on wire gauze on a retort-stand ring about six inches above a small flame to dry slowly (Fig. 81).

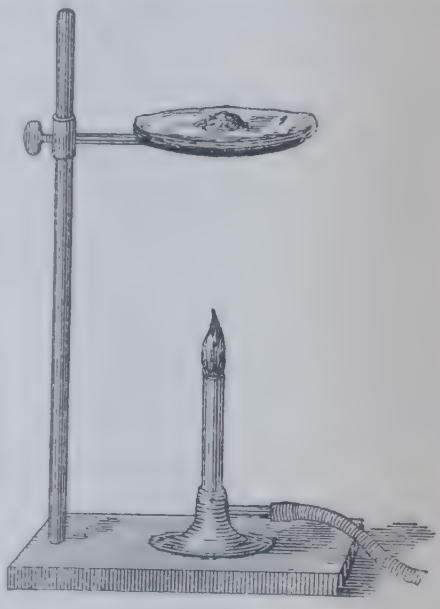

FIG. 81.

But even when both constituents of a mixture are soluble in water, it is possible to separate them, provided they are soluble in different degrees.

EXPT. 9. **To separate Nitre from Common Salt—Two Substances which are both Soluble in Water.**—Although both substances are soluble, they are not equally soluble. Thus 100 grams of water are able to dissolve

	at 0° C.	at 100° C.
Nitre	12 gm.	247 gm.
Salt	36 ,,	38 ,,

Salt is almost equally soluble in hot and in cold water. Nitre is less soluble than salt in cold water, but far more soluble in

hot water. What method would you suggest for separating nitre from a mixture of nitre and salt? Weigh out 30 grams of nitre and 30 grams of salt, and grind them together in a mortar. Measure 40 c.c. of water into a porcelain basin, heat it, and then slide in the mixture. Continue to heat with a small flame, and stir with a glass rod; a white granular substance refuses to dissolve (what is it?). Meantime heat a funnel fitted with a cone of filter-paper by pouring boiling water through it several times, and then at once filter the boiling solution, catching the filtrate in a small beaker. Cool the beaker by floating it in a tumbler containing cold water; long needles of nitre crystallise out. When cold, pour off the mother-liquor as completely as possible, and dry the crystals on filter-paper. These crystals may, however, still contain a small quantity of salt. Redissolve them, therefore, in the least possible quantity of boiling water. Again cool; much of the nitre again crystallises, but there is ample water to retain all the salt in solution. Dry the crystals on filter-paper. Find what weight of nitre has been recovered.

WATER OF CRYSTALLISATION

Try the following experiments:—

EXPT. 10. **To observe the Effect of Heat on Crystals of Blue Vitriol.**—Heat a few small crystals of blue vitriol in a small dry test-tube, holding the tube nearly horizontally. Observe that a heavy dew is deposited on the upper part of the tube, and at the same time the blue crystals become ashen grey and non-crystalline or *amorphous*. Steam also escapes from the mouth of the tube. Allow to cool: then add 4 or 5 drops of water to the residue in the tube. What do you observe?

It appears that heat drives water out of the blue crystals, and at the same time destroys their colour and crystalline form. The addition of water to the amorphous residue restores the blue colour. Can the crystalline form also be restored?

EXPT. 11. **To attempt to recover Crystals by the Addition of Water to the Amorphous Residue.**—Warm the test-tube containing the residue, to which a few drops of water have been added. Pour the hot deep blue solution into a watch glass, and allow to cool. Crystals of blue vitriol are formed.

Then it seems that the crystalline form has been recovered, because water has been added. Moreover, the vapour which was given off when the crystals were heated must have been water vapour. Water which can be driven off by heat in this way from crystals is called **water of crystallisation**, and generally the crystals can be re-formed by the addition of water. We shall find that there are many solid crystalline substances, *e.g.* soda crystals, Epsom salts, which contain such water, though there are others, *e.g.* nitre, calcspar, which do not contain any water.

EXPT. 12. **To find the Percentage of Water of Crystallisation in Blue Vitriol.**—Weigh a small dry porcelain crucible and lid. Add 1 to 2 grams of finely powdered blue vitriol. Weigh exactly. Place the crucible on a sand-bath with the lid on (slightly tilted), pack it round with sand (Fig. 82) Heat it for twenty minutes with the flame about three-quarters of its full height. Carefully remove the crucible from the sand-bath so as to cool. Weigh. Re-heat for ten minutes and again weigh, and so on until there is no further loss in weight. Record thus :—

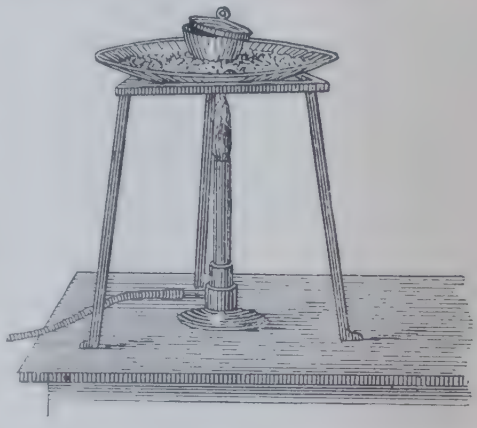

FIG. 82.

Weight of crucible	= 12·35 gm.
Weight of crucible + blue vitriol	= 14·03 ,,
∴ Weight of blue vitriol	= 1·68 ,,
Weight of crucible + residue, heated 20 mins.	= 13·42 ,,
Weight of crucible + residue, heated 10 mins. more ...	= 13·42 ,,
∴ Loss in weight	= 0·61 ,,
Hence, 1·68 gm. blue vitriol lost	0·61 ,,
∴ 100 gm. blue vitriol lost	$= \frac{0{\cdot}61}{1{\cdot}68} \times 100$,,
	36·08 gm.

Do you find the same loss?

Problems and Exercise

Crystallisation

1. Prepare crystals of the following substances:—Rochelle salt, chrome-alum, copper sulphate, ferrous ammonium sulphate, zinc sulphate.

2. Prepare crystals of salt by pouring a saturated solution of salt into a soup plate, and putting it where it may gradually evaporate undisturbed and free from dust.

3. Pour a dilute warm solution of sal ammoniac over a warm sheet of glass. Allow to cool. Examine the crystals which form with a microscope, and draw their shape.

4. Find what happens when crystals of chrome-alum are placed in a crystallising solution of common alum (potash alum).

5. Make a mixture of chrome-alum and potash alum solutions. Crystallise.

6. Make mixtures of the following substances, and then separate samples of the constituents by means of water:—

Chalk and common salt.
Nitre and charcoal.
Sand and sugar.

7. From the muddy liquid, containing a soluble and an insoluble substance,[1] separate the whole of the latter, clean and dry.

8. From the given powder,[2] containing a soluble and an insoluble substance, separate the whole of the latter, clean and dry.

9. Find the percentage of insoluble matter in the given mixture.[2]

10. Separate from gunpowder the part which is soluble in water.

11. Mix thoroughly 40 grams of chlorate of potash with (*a*) 40 grams potassium chloride, (*b*) 40 grams bichromate of potash. Refer to a text-book for their solubility curves. Then separate by crystallisation a sample of pure chlorate of potash from each mixture.

12. Find the percentage weight of water of crystallisation in 100 grams of the following substances: A, B, C, D, E.[3]

[1] *E.g.* $KClO_3$ and MnO_2; $KMnO_4$ and MgO.
[2] *E.g.* $K_2Cr_2O_7$ and $CaCO_3$; $KClO_3$ and MnO_2; $NaCl$ and SiO_2.
[3] The following are suitable substances: $CuSO_4 . 5H_2O$, $ZnSO_4 . 7H_2O$, $Na_2CO_3 . 10H_2O$, $MgSO_4 . 7H_2O$, $BaCl_2 . 2H_2O$.

CHAPTER XV

RAW MATERIALS

The Earth is the Chemist's Storehouse.—The most careless observer cannot fail to conclude that the soil and rocks which form the crust of the earth are not everywhere of the same nature. The purple hue of Welsh slate, the transparency of Derbyshire spar, the sombre gray of Scotch granite, or the yellow sands of the sea-shore point to differences of material. The early alchemists understood this, and stored the different minerals and ores they found in the crust of the earth upon the shelves of their laboratories. They made experiments with them in their crucibles and alembics, and succeeded in producing many new substances.

So, too, for us the earth is a storehouse, and we have but to search it for raw material. We may obtain common salt from Northwich, chalk from Dover, iron ore from Spain, or saltpetre from India. With these and other native or raw materials we may then make all manner of experiments and hope to prepare new substances.

It will be wise, therefore, in the first place to obtain some of these materials, and proceed to examine them, noting points of resemblance and difference. We must carefully observe their appearance, forms, and their colours ; their taste and the action of water upon them may also be tried, and it will perhaps be useful to try the effect of heat upon each ; then when a mineral is brought to us to be named we shall often be able to identify it.

1. Common Salt

Common salt is always present in sea water along with other substances, and for ages men have obtained it from the waters of the oceans. If sea water be evaporated, cubical crystals of common salt separate out. These must be purified by re-crystallisation. In warm climates the evaporation is often done in great basins on the sea-shore, and the crystals of salt are raked off. "Bay salt," which may be purchased through a chemist, shows the cubical crystals.

Common salt also occurs as huge deposits of rock-salt in various strata of the earth. The chief of these in England is at Northwich in Cheshire, where there are two great beds of rock-salt, together nearly 60 feet in thickness. It is probable that this rock-salt was deposited ages ago by the evaporation of sea water. A common method of bringing the salt to the surface from these buried stores, is to bore down to them and to pour water into the boring. After lying in contact with the salt beds, the water becomes thoroughly saturated. The rich brine is then pumped up to the surface and evaporated. One hundred parts by weight of brine will yield about 23·5 parts of salt.

Expt. **Is Common Salt changed by Heat?**—Place a little common salt in a small dry test-tube. Heat it with a Bunsen burner. The salt is scarcely changed in appearance. If the salt be pure, no vapour is given off. Frequently common salt is moist; if so, a dew of water will be deposited on the glass. A crackling noise is also heard as the crystals break and fly into smaller crystals. If the flame is a good one, the salt may at last melt.

2. Chalk

Chalk is found right across England from Flamborough Head, along the Wolds of Yorkshire and Lincolnshire, past Newmarket and the Gog Magog Hills in Cambridgeshire, to Salisbury Plain. Thence it stretches in the North and South Downs to Dover and Beachy Head. Chalk has been shown to consist largely of the broken fragments of tiny shells or foraminifera, such as are now forming the gray ooze or soft

bed at the bottom of the Atlantic, on which the telegraph cables rest (see Fig. 83).

EXPT. **To find the Tiny Shells in Chalk.**—With a camel's hair brush dust the surface of a lump of chalk into a watch-glass containing some water. Carefully pour off the water from the watch-glass, and then fill it up again with water. In this way wash the dust again and again with water. The heavier particles sink to the bottom, the lighter are swept away. Examine the residue with the microscope. Real chalk must be

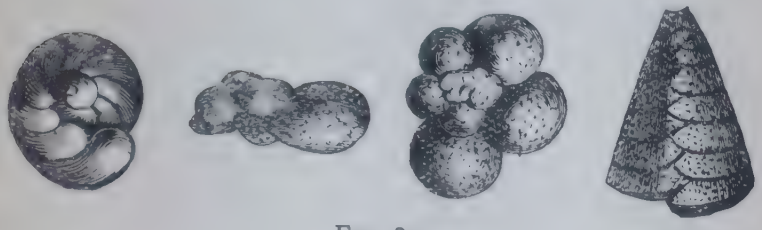

FIG. 83.

used for this experiment, not the soft "chalk" used on blackboards.

For further information the student is recommended to read Huxley's "Physiography" on this subject.

EXPT. **The Effect of Heat on Chalk: Formation of Quicklime.**—Heat a little powdered chalk on platinum foil in a Bunsen flame. It does not change in appearance. Remove it, and allow it to cool. Taste a very small portion. It is very caustic, and not earthy like chalk.

It has long been known that when chalk or limestone is exposed to a violent heat, it is converted into this caustic substance, called **quicklime**.

In chalk and limestone districts you may have seen the kilns in which quicklime is obtained by heating a mixture of chalk or limestone with a little coal.

Limestone is the general name given to many stones which when burnt in a kiln yield lime. To include all such stones we ought to call *chalk* and *marble* limestones. A large part of the Pennine Range consists of hard dark blue-gray stone streaked with white, called "mountain limestone." It contains numerous fossils of shells and corals, suggesting that it has once been mud at the bottom of a coral sea. It is found from Cheviot past

Crossfell and Ingleborough to the Peak. The streams which flow over it, like those in chalk districts, are hard. Their waters dissolve the rock into fantastic caverns.

Another sort of limestone called *Dolomite* or *Magnesian Limestone* extends in a narrow band from the coast of Durham past Knaresborough and Pontefract in Yorkshire, to Nottingham.

3. Soda Crystals

Up to the beginning of the last century the sea-weeds cast up on the coasts of Normandy, Scotland, and Ireland were the only source of soda. The seaweeds were burnt, and their ashes—called *kelp*—were dissolved in hot water. On evaporating the liquor, crystals of soda were obtained. It is now made from common salt in very large quantities and by more complicated processes.

(1) Examine and describe some soda crystals. They are of irregular shape, white, and semi-transparent.

(2) Taste a crystal. It has a peculiar brackish "**alkaline**" taste.

(3) Shake a few powdered soda crystals with cold water in a test-tube. They dissolve extremely easily. Place a drop or two of the solution on the palm of the hand, and rub with a finger. It produces a soap-like sensation. Hence, the substance is called **washing-soda**, because it is found to have the same cleansing action as soap, but it is sometimes simply called *soda*.

(4) Find the effect of heat on soda crystals. Proceed exactly as in Expts. 10, 11, and 12, p. 141. Do you not find that the crystals seem to contain a good deal of water of crystallisation? How much? Nearly 63 per cent.

4. Potash[1]

When land plants are burnt, an ash is left which contains a substance very similar to soda. This used to be done in large pots, and so the ash was called *pot-ash*. The ash was then treated with hot water, and the solution so obtained was filtered from any undissolved matter. On evaporating the solution, a

[1] Potassium carbonate

brown mass of crude potash was obtained. To purify it, it was heated strongly, and then re-treated with water.

(1) Examine and describe potash. It is a white granular non-crystalline powder.

(2) Taste a little. It has the same alkaline taste as soda.

(3) Try the action of cold water on it. It is extremely soluble, and the solution produces the same soap-like sensation.

(4) Find the effect of heat on potash in a small dry test-tube. Does it melt? Is vapour given off? No. Potash differs, in fact, chiefly from soda in not containing water of crystallisation.

5. Saltpetre or Nitre

In certain parts of India, the soil, after the rainy season, becomes covered with small crystals of a different salt from any referred to so far in this chapter. This "salt of the rock" was called *sal petræ* by the alchemist Geber, and is known at the present day as *saltpetre* or *nitre*. To separate this substance, the soil is scraped and treated with water, in which the nitre readily dissolves, and on the evaporation of this solution the nitre crystallises out.

(1) Examine and describe carefully the appearance of nitre. Make a drawing of a good crystal. It is a long 6-sided prism, with wedge-shaped ends.

(2) Taste a small crystal of nitre. It has a bitter cooling taste.

(3) What has been found to be the action of water on nitre?

(4) **The Effect of Heat on Nitre.**—Place about 1 gram of nitre crystals in a dry test-tube. Heat in a Bunsen flame. The crystals melt, and the liquid seems to boil, and a number of little bubbles are evolved. Allow to cool. It solidifies, forming a white opaque cake, which is crystalline unless the nitre was heated a long time.

6. Sulphur

Sulphur or brimstone (burning stone) has been known since very early times, for it is often present in large quantities in volcanic districts. This is particularly the case in the island of

Sicily, where it lies embedded in the rocks. Examine a specimen of native sulphur.

Sulphur, we have already found, is readily melted, and it will itself burn. In order, therefore, to separate the sulphur from impurities, it is only necessary to pile a heap of ore on the slope of a hill, and kindle the heap at the bottom; the heat from the sulphur which burns is sufficient to melt the rest, and this trickles down and is run into moulds, where it solidifies.

EXPT. **To prepare "Flowers of Sulphur."**—Fix a hard glass test-tube by means of a cork in the wide mouth of a lamp

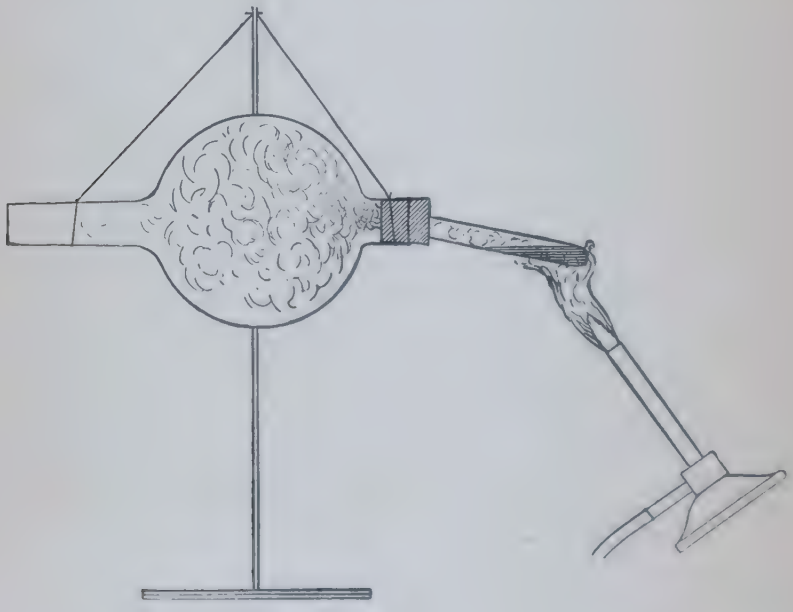

FIG. 84.

glass, as drawn (Fig. 84). Place a few lumps of sulphur in the test-tube. Sling the whole by a string from a ring of a retort-stand. Heat the sulphur gently. It readily melts, forming a mobile pale yellow liquid (at 114° C.), and then it quickly darkens in colour, and becomes as thick as treacle: with more heat it again becomes a thin liquid, but it remains very dark in colour, and presently boils: the temperature is then 448° C. The vapour passes into the globe of the lamp-glass; much of it is condensed as a yellowish powder, but as the glass walls become hot, some may collect at the bottom as a liquid, which

quickly solidifies. The fine light yellow powder is called "**flowers of sulphur**."

This process is carried out on the large scale by boiling sulphur in an iron retort, G, and sending the vapours into a great brick chamber, A, 200 cubic metres in capacity (Fig. 85). After a time the walls of the chamber become so hot that the vapour is not condensed as flowers of sulphur, but as a liquid, which collects on the floor. It is then run out and cast in wooden moulds in the form of sticks. This is called **roll sulphur**. Thus the process is a distillation, and all impurities are left behind.

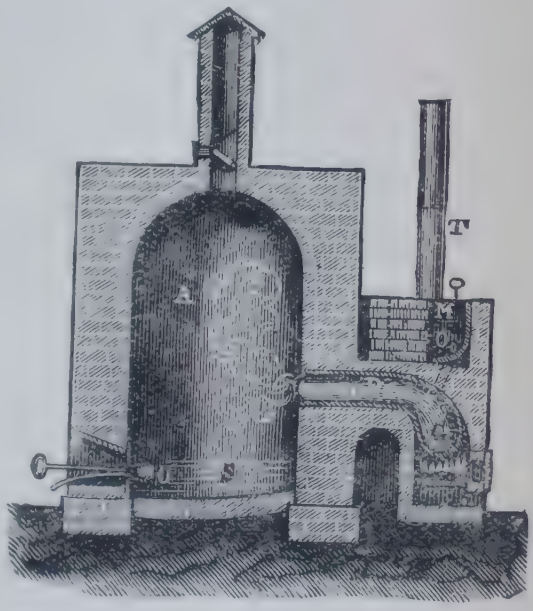

FIG. 85.

Make the following simple experiments with some roll sulphur.

(1) Examine and describe a piece. Crush it gently with a pestle in a mortar; it is very brittle. Notice a fresh surface, and observe that the lump is a matted mass of crystals.

(2) Heat a few small pieces in a dry test-tube, and notice the changes it undergoes.

(3) Invert a crucible lid on a pipe-clay triangle. Place a small piece of sulphur on the lid, and heat the lid until the sulphur catches fire. Notice the blue lambent flame, and the suffocating "sulphurous" odour.

(4) Find whether sulphur is at all soluble in water. It is quite insoluble.[1]

The effect of cooling molten sulphur rapidly by pouring it into water is remarkable:—

[1] Roll sulphur is, however, soluble in a liquid called carbon bisulphide. Shake some powdered roll sulphur with a little carbon bisulphide in a dry test-tube. The sulphur is dissolved. Pour the solution into a watch-glass, and let it evaporate in a fume-cupboard. Crystals of sulphur are recovered.

EXPT. **To prepare "Plastic Sulphur."**—Fill a large hard glass test-tube about one-third full with powdered roll sulphur. Fill a beaker with cold water, and invert a funnel in it. Heat the sulphur cautiously, until it is a very dark mobile liquid, and then pour it in a thin stream into the water, round and round the funnel. Lift out of the water the funnel with the sulphur, and notice that it is no longer crystalline and hard, but soft and elastic like india-rubber.

That this curious substance is still *sulphur*, though in a different form, is shown by the fact that in a few days it becomes hard and opaque, and like ordinary brittle crystalline sulphur. This soft variety of sulphur is called "**plastic sulphur.**"

7. IRON PYRITES

Iron pyrites occurs very widely in many different places. When pure it has a brass-like colour. Small quantities of it may frequently be found in lumps of coal.

Examine a piece carefully. Perfect crystals of iron pyrites take the forms of the regular solids; the cube and the octahedron are common forms.

EXPT. **The Effect of Heat on Iron Pyrites.**—Powder a small quantity, and heat it in a dry test-tube. A sulphurous vapour is given off, and a yellow deposit of sulphur is formed in the cooler part of the tube, while the mineral acquires a deep red colour. If the yellow deposit is heated, it melts and moves up the tube.

8. GREEN VITRIOL[1]

When iron pyrites is long exposed to the air and moisture, a greenish liquor is obtained. If the liquor is evaporated, large green glass-like crystals form. These were known to the alchemists under the name of *green vitriol*.[2]

At the present time large quantities of green vitriol are obtained in this way from the pyrites occurring in the coal of South Lancashire.

1. Examine and describe some green vitriol.

2. Taste a small crystal of green vitriol. It has a sweet and inky taste.

[1] Sulphate of iron. [2] Latin *vitrum*, glass.

3. Try the action of cold water on it. It is readily soluble.

4. **The Effect of Heat on Green Vitriol.**—Place a few crystals in a dry test-tube. Heat cautiously in a flame. It melts, boils, and a dew of water is deposited on the cool upper part of the tube. Soon it becomes dry, and a hard white mass is left. Then heat strongly. A sulphurous vapour is given off, and the mass becomes red, where it is in contact with the glass. This red substance is used as *rouge*.

9. Blue Vitriol

The drainage water from copper mines has frequently a blue colour. If such water is partially evaporated, and the liquid is then allowed to cool, fine blue crystals are obtained. These were known to the alchemists as *blue vitriol*.

1. Examine and describe some blue vitriol.
2. Do not taste blue vitriol. It is poisonous.
3. Try the action of cold water upon it. It is fairly soluble in cold water, and much more so in hot water.
4. What has been found to be the effect of heat upon crystals of blue vitriol? See p. 141.

10. Sal-Ammoniac[3]

A remarkable salt was prepared long ago in Egypt from camel's dung. When the dung was burnt, a heavy soot was deposited, from which the salt called *sal-ammoniac* was made. In those days, just as at the present time, it was widely used as a medicine.

Later, it was prepared from the soot obtained when such animal refuse as hoofs, horns, and hair were heated.

1. Examine and describe the appearance of sal-ammoniac. It consists of minute white crystals.
2. Taste a small quantity. It is sharp and cooling.
3. Try the action of water upon it. It is very soluble.
4. **The Effect of Heat on Sal-Ammoniac.**—Place a small amount of sal-ammoniac in a dry test-tube. Heat gently. White fumes are given off. Finally the whole of the sal-ammoniac disappears from the bottom of the tube; part has escaped into the air, but much is found on the cooler part of

[3] Ammonium chloride.

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the tube. Examine some of this deposit: taste it, and try the action of water upon it It behaves in exactly the same way as the original substance.

This is an example of what is called **sublimation**. The sal-ammoniac which has been re-formed on the side of the tube by the condensation of the vapour of the original sal-ammoniac is called a **sublimate**, and the original sal-ammoniac is said to have **sublimed**.

Problems and Exercises

Raw Materials

1. Describe fully a piece of slate.
2. Describe fully a lump of granite.
3. Describe fully the appearance of the substances, A, B, C, D, E, F.[1]
4. Find the effect of heat on the substances, G, H, I, J, K, L.[2]
5. Identify the substances, M, N, O, P, Q, R.[3]
6. Find whether any water is contained in the substances, V, W, X, Y, Z.[4]
7. You are given a powder.[2] Describe
   (1) its appearance.
   (2) action of cold water upon it.
   (3) effect of heat in a dry test-tube.

[1] *Note for the Teacher* – The following are suitable substances, *e.g.* HgO, HgI, Pb_3O_4, NH_4NO_3, HgS, Sb_2S_3.

[2] *e.g.* $NaHCO_3$, $Pb(NO_3)_2$, $ZnSO_4$, $7H_2O$, $KClO_3$, camphor, iodine, cobalt chloride, Rochelle salt.

[3] *e.g.* NaCl, KNO_3, $CuSO_4$, CaO, NH_4Cl, Na_2CO_3.

[4] *e.g.* $Na_2SO_4 . 10H_2O$, Rock Salt, KNO_3, $K_2Cr_2O_7$, $ZnSO_4 . 7H_2O$.

CHAPTER XVI

THE PREPARATION OF THE COMMON ACIDS AND ALKALIES

THE Alchemists made numberless experiments with the raw or native materials which they found ready for them in the crust of the earth, and learnt how to prepare many new substances. Amongst these none have proved of greater importance and have found more frequent use than certain liquids which have strong acid or "alkaline" tastes.

We also must now learn how to prepare these liquids for future use, and how to distinguish one from another, by such simple tests as we can discover.

A. THE COMMON ACIDS

1. Oil of Vitriol[1]

EXPT. 1. **Preparation of Vitriolic Acid from Green Vitriol.**[2]—Weigh out about 10 grams of green vitriol crystals into an iron dish or porcelain basin. Heat over a flame, stirring with a glass rod until all water is expelled and a brittle white mass is left. Break the mass into small lumps (but do not powder it). Introduce these into a small retort, *a* (Fig. 86). Support the retort so that it can be heated by a naked flame. Let the neck dip into a little water in a test-tube, *b*. Heat the retort at first with a small flame, and finally with the full flame for at least fifteen minutes. Remove the test-tube. Then turn the flame down gradually, so that the retort may cool slowly, but the experiment is almost certain to result in the destruction of the retort.

[1] Sulphuric acid.

[2] Lecture Table Experiment.

Is the water in the test-tube still merely water? Pour two or three drops into another test-tube. Taste these. The liquid has a very acid taste. This acid liquid, thus prepared from green vitriol, was called by the early chemists **vitriolic acid**. A better method of preparation has been discovered since the time of the alchemists, but this is not the place to describe it.

Transfer the residue in the retort *a* on to a sheet of white paper. Observe that those portions which were most strongly

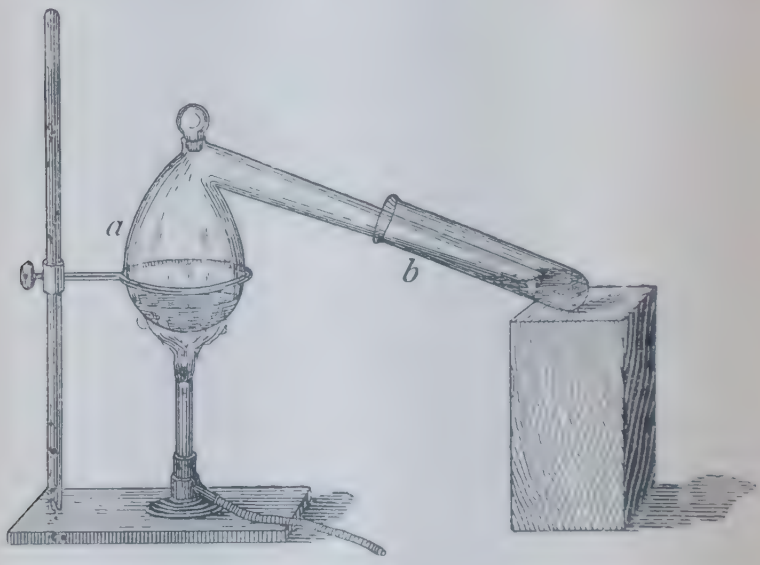

FIG. 86.

heated have acquired a deep red colour. This substance when mixed with oil is used largely as a red paint.

EXPT. 2. **Vitriolic Acid when heated gives off Dense White Fumes.**—Pour some of the acid prepared in Expt. 1 into a small porcelain crucible. Heat it over a small flame. It boils down to a small bulk giving off steam. Then the fumes become thick and white, and possess a pungent odour, and the remaining liquid has an oily appearance. At once remove the flame, and take the crucible to a fume-cupboard. On this account the acid is frequently called **oil of vitriol**. Another name for it, much used now, is **sulphuric acid**.

EXPT. 3. **The Mixing of Oil of Vitriol and Water.**—Three-quarters fill a test-tube with water. Mark the level of

the water on the side of the glass. Measure into another test-tube about one-eighth of this volume of acid. Pour some of the acid into the water in a thin stream. It sinks to the bottom ; so the acid is heavier than water. Stir the water round with a glass rod. Add the rest of the acid in the same way. Grasp the test-tube in the hand. The mixture has become quite hot, and steams.

Oil of vitriol when mixed with water becomes sometimes hotter than the boiling point of water. What would happen if a little water were poured on to strong oil of vitriol? Being lighter, the water would float on the surface. The liquids would mix at the bounding surface, and become very hot, hot enough to turn some of the water to steam, and throw the acid about with explosive violence. Hence, *never pour water on to strong oil of vitriol.*

EXPT. 4. **To show the Destructive Action of Oil of Vitriol.**—(*a*) Write a word with some dilute oil of vitriol with a glass rod on a piece of *paper*. Dry the paper thoroughly by gently wafting it above a Bunsen flame. The paper is charred where the acid was traced.

(*b*) Place about 10 grams of lump *sugar* in a tall beaker or jar. Add about 10 c.c. of hot water. Place the beaker on a plate or dish. Then add at once about 10 c.c. of strong oil of vitriol. The mixture blackens at once, and froths up.

(*c*) Mix 1 or 2 c.c. of oil of vitriol with about 10 c.c. of water. Pour the mixture on a small piece of *zinc* in a beaker or basin. There is a violent effervescence. The zinc is dissolved.

Note on Sulphuric Acid.—(1) Should any acid be spilt, throw on some quicklime or powdered chalk, and mop up with an old duster. (2) Never let the acid boil : the fumes are very disagreeable.

After the discovery of oil of vitriol the alchemists tried the effect of heating it with many substances, with the result that they were led to discover other acid liquids, as will be described in the course of the following pages.

2. Spirit of Salt [1]

Another acid liquid was obtained by the alchemists by heating common salt with certain substances. Glauber, who

[1] Hydrochloric Acid or Muriatic Acid.

lived at the beginning of the seventeenth century, first prepared it from common salt and oil of vitriol, and it is always made at the present time from these.

EXPT. 5. **Preparation of Spirit of Salt.**—Weigh out about 30 grams common salt. Introduce into a litre flask. Attach a cork fitted with a thistle-funnel, and a delivery tube bent twice at right angles. Attach an inverted funnel, *a*, to the end of the delivery tube by a small piece of rubber tubing (Fig. 87). Support the flask upon wire gauze on the ring of a retort-stand, and place the funnel within a tumbler, *b*. Add enough water to the tumbler to cover the mouth of the funnel.

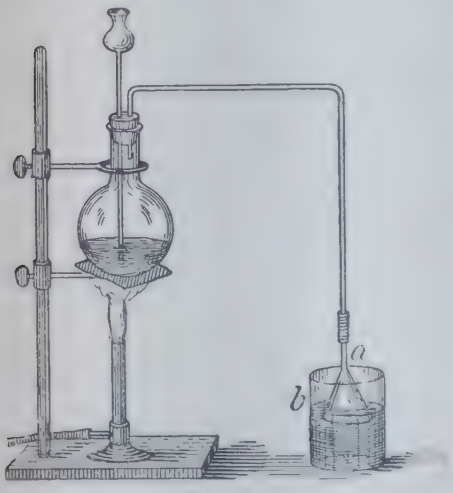

FIG. 87.

Measure out 30 c.c. strong oil of vitriol. Pour it little by little down the thistle-funnel on to the salt. Fumes are given off, and bubbles are forced through the water. As soon as all the acid has been added, heat the flask *gently*. Soon the bubbles cease to come through the water.[1] After a time the water in the tumbler begins to fume strongly. As soon as this is the case, remove the tumbler and also the flame. Dilute the contents of the flask, and pour away.

Has anything passed into the water in the tumbler? It seems probable from the way in which it fumes. Dip a glass rod into the water, and place a very small drop upon the tongue. It is very acid, and has a different taste from oil of vitriol. Where has the acid come from? The alchemists thought that it was a spirit which the oil of vitriol had driven out of the salt. They therefore called it **spirit of salt**. Later it was called **muriatic acid**, from the Latin *muria*, brine; but the fuming acid liquid is now called **hydrochloric acid**.

[1] Care must be taken that the water in the tumbler does not rush back along the tube into the flask. If it begins to do so, at once heat more strongly, or remove the tumbler.

EXPT. 6. **Is Spirit of Salt Volatile?**—Pour about 5 c.c. of the acid into a porcelain dish. Heat it gently over a flame. Acid vapours are given off it, and it is soon completely evaporated.

Try whether spirit of salt will attack paper and zinc.

3. Spirit of Nitre

The alchemist Geber, who lived about the twelfth century, described an acid liquid obtained by heating nitre with green vitriol. Since the acid liquid was obtained by means of nitre,

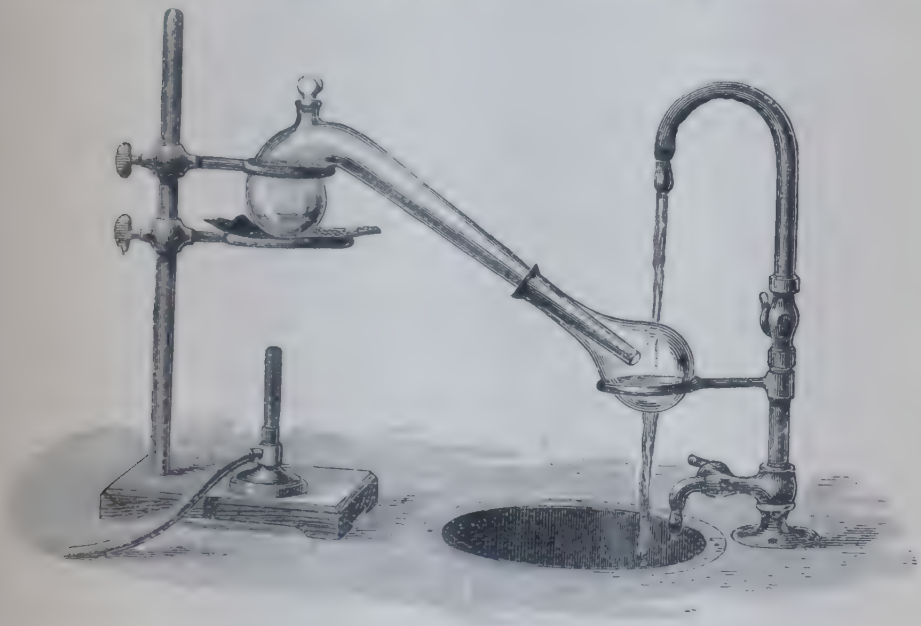

FIG. 88.

it used to be called **spirit of nitre.** It is now usually termed **nitric acid.**

Although spirit of nitre can readily be obtained by Geber's method, it is now usually prepared by heating nitre with oil of vitriol in place of green vitriol. Glauber first prepared it in this way, and he called it "spiritus nitri fumans Glauberi" or "Glauber's fuming spirit of nitre."

EXPT. 7. **Preparation of Nitric Acid from Nitre and Oil of Vitriol.**—Weigh out 20 grams of nitre, and measure out 30 c.c. strong oil of vitriol. Support a small retort on wire gauze, as shown in Fig. 88. Pass the neck into a small dry

flask. Support the flask under a tap or in a trough, and keep it cool with running water.

Slide the nitre into the retort, and pour the oil of vitriol in through a funnel, taking care that neither gets down the long tube of the retort. Replace the stopper of the retort. Heat gently. The nitre melts, and the liquid begins to boil. A ruddy vapour rises, and drops form and run down the tube. A yellowish liquid collects in the flask. Keep the liquid just boiling, but do not let the temperature rise unnecessarily. As soon as the liquid in the retort begins to thicken, remove the flame.

Pour the remaining liquid from the retort into an empty evaporating basin. Put it aside. On cooling it hardens into a solid mass. If this had been left in the retort, it would have been rather difficult to get it out. Wash out the retort when cold.

Can nitric acid be evaporated just as spirit of salt (hydrochloric acid)?

EXPT. 8. **Is Nitric Acid Volatile?—Tests for Nitric Acid.**—(1) Pour about 5 c.c. of the liquid into a porcelain dish. Heat gently until the liquid is evaporated. Vapours are evolved possessing a peculiar odour quite unlike the fumes of oil of vitriol or of spirit of salt.

(2) Place a small drop of the liquid upon a finger nail, and another upon the skin of the hand. Wash them off after a few seconds. Bright yellow stains are left.

(3) Place a small piece of copper in the liquid. Red fumes are at once evolved, and the liquid becomes green or blue.

Show that oil of vitriol and spirit of salt do not attack copper in this way. On account of the corrosive action of nitric acid it used frequently to be called **aqua fortis** (*i.e.* strong water).

Try whether nitric acid will attack paper and zinc.

4. *Vinegar or Acetic Acid

Acetic acid is another acid which is frequently used, and therefore may be mentioned here, although it is not prepared from any materials which are found native in the rocks or crust of the earth. It was the only acid known to the ancients.

When beer or a weak wine such as claret is left exposed to

the air, it quickly becomes sour. This is because the alcohol has been changed into the acid called acetic acid or vinegar.

The best way to convert weak wine into vinegar is the German "quick-vinegar process." A large cask with perforated sides is filled with beech-wood shavings, which are moistened with a little vinegar. Weak wine is then poured in at the top, and it trickles slowly through the shavings, and the liquid which collects at the bottom is again and again poured down until all the alcohol is changed into vinegar. A kind of fungus called *mycoderma aceti* is found on the shavings, and it is this which in the presence of air effects the change.

EXPT. 9. **Is Acetic Acid Volatile?**—Heat about 5 c.c. of the acid in a basin. Observe that the vapour has a very pungent odour, which is quite different from that from any of the other acids. Little or no residue is left.

Try whether acetic acid will attack zinc, copper, or paper.

(B.) SOME COMMON ALKALIES

We must now learn a little about some substances very different in character from the acids, which the alchemists also knew how to prepare.

1. Caustic Soda

It was found long ago that if a solution of soda crystals was heated with lime the solution changed in character and acquired a very caustic nature. You will have an opportunity of trying this yourself later (see Volume II.).

If such a solution is allowed to stand till clear, and is then evaporated, a white non-crystalline solid is obtained. A piece of this solid or of the strong solution will quickly cause a blister when placed upon the skin. It also destroys vegetable fibre, *e.g.* wood or filter-paper. On this account it is called **caustic soda**, and must be handled with care.

Obtain a piece of solid caustic soda. Examine it, and see how readily it dissolves in water.

Dilute with a large amount of water a little of a solution of caustic soda, such as is in a bottle on your shelf. Taste it, and notice its soapy "**alkaline**" nature. Wet your fingers with the solution: they feel soapy and slippery. Next time your

hands are greasy with oil, try washing them in a dilute caustic soda solution. They will become beautifully clean. A stronger soda solution would remove the skin as well as the dirt.

2. Caustic Potash

Another caustic solution is formed if a solution of potash is heated with lime. If the clear caustic solution is evaporated, a white solid is obtained which is almost exactly similar in appearance to caustic soda. Moreover its solution in water has a very similar caustic and alkaline nature.

This substance is called *caustic potash*. It is usually sold in white sticks.

3. Volatile Spirit of Sal-Ammoniac or Ammonia

EXPT. 10. **Preparation of Spirit of Sal-Ammoniac or Ammonia.**—Weigh out 10 grams sal-ammoniac. Powder it.

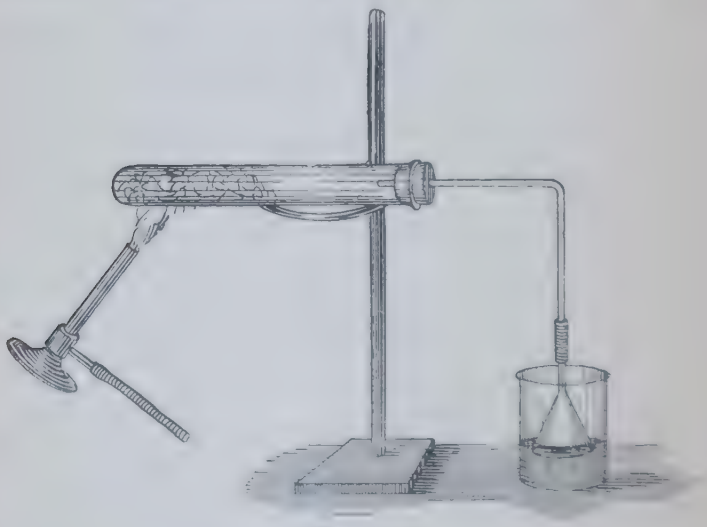

FIG. 80.

Dry it by heating it gently in a basin on a sand-bath. Allow it to cool. Meantime weigh out 15 grams quicklime, and powder it. Mix the two together. Put the mixture in a large dry test-tube. Fit the tube with a cork and a delivery tube bent once at right angles. Attach an inverted funnel to the delivery tube. Support the test-tube on the ring of a stand, so that the funnel dips into a tumbler half filled with water,

with its mouth *just below* the surface of the water, as shown in Fig. 89.

Heat cautiously by wafting the flame about under the tube. At first bubbles of air are forced up through the water, and then these cease. Heat five minutes longer. Take care that the water in the beaker is not sucked into the test-tube. Then remove the beaker, and afterwards the flame.

Observe the pungent and peculiar odour the water in the beaker has acquired. As this liquid was acquired by heating sal-ammoniac with lime, the early chemists called it the **volatile spirit of sal-ammoniac.** It is now usually called **ammonia.**

Ammonia can also be prepared by heating shavings of horn. Indeed, this method of preparation was at one time so commonly used that the pungent liquid came to be known as "spirits of hartshorn," a name which is still common for it.

EXPT. 11. **Tests for Ammonia.**—(1) Pour a little of the solution into a porcelain crucible. Heat over a small flame. It is quickly evaporated, and the vapours possess the characteristic odour already observed.

(2) Dip a glass rod in strong hydrochloric acid. Hold it in the vapours arising from the solution. Dense white fumes are formed.

Repeat with a rod dipped in sulphuric acid and in nitric acid.

4. Lime Water

EXPT. 12. **Slaked Lime is somewhat Soluble in Water.**—Half fill a test-tube with water. Add some powdered slaked lime. Shake. The liquid becomes white and milky. The lime has not all dissolved. Perhaps some of it has dissolved. Allow the test-tube to stand. The white mud, which is only in suspension, begins to settle. In time it will all settle, and leave the liquid quite clear. To save time filter the liquid. Evaporate a drop on platinum foil. A white stain is left, showing that some of the slaked lime was dissolved.

Slaked lime, then, is slightly soluble in water. The solution is called **lime water.** Observe its alkaline taste. Lime water is generally prepared by shaking quicklime with water, and allow-

ing it to settle. The shaking is repeated to strengthen the solution. The clear liquid is then decanted or syphoned off.

The acid and alkaline solutions we have been studying are commonly spoken of as *acids* and *alkalies*.

Problems and Exercises

The Common Acids and Alkalies

1. Find the effect of evaporating the liquids, A, B, C, D, E,[1] in a porcelain basin.

2. You are given samples of the following pairs of liquids. Distinguish the members of each pair.

Oil of vitriol and nitric acid.
Hydrochloric acid and nitric acid.
Oil of vitriol and acetic acid.
Acetic acid and ammonia.

3. Find the effect of heating 1 part of nitre with 2 parts of green vitriol in a small retort the neck of which dips into a little water in a flask.

4. Find the effect of heating sal-ammoniac with oil of vitriol, as in Expt. 5.

5. Find the effect of heating Chile saltpetre with oil of vitriol, as in Expt. 7.

[1] *Note for Teacher.*—*e.g.* NH_4OH, $Ca(OH)_2$, H_2SO_4, HCl, HNO_3, $CH_3 . COOH$.

CHAPTER XVII

THE ACTION OF ACIDS AND ALKALIES UPON LITMUS—THE FORMATION OF SALTS

CHEMISTS found 200 years ago that the acid and alkaline substances we have just learnt to prepare had a curious effect on certain vegetable dyes. Of these dyes *litmus* is one of those most readily acted on, and it is therefore very suitable for our experiments. It is a solid blue substance, which is extracted from certain kinds of lichen.

EXPT. 1. **To make an Infusion of Litmus.**—Weigh out about 5 grams of solid litmus. Grind it up in a mortar, and transfer into a flask. Half fill the flask with cold water, and shake well. Allow the powder to settle, and then pour off and throw away the liquor. In this way certain impurities are washed out of the litmus.

To the residue add 100 c.c. of water, and heat the flask nearly to boiling for a few minutes. Allow to settle[1] (if possible for some hours in a warm place). Decant off the intensely blue solution from the sediment, and place the solution in a bottle.

EXPT. 2. **To find the Effect of Acid and Alkaline Liquids on the Blue Litmus Solution.**—Let us write down the names of the acid and alkaline liquids we have on our shelves :—

Acid liquids.	*Alkaline liquids.*
Sulphuric acid.	Caustic soda.
Hydrochloric acid.	Caustic potash.
Nitric acid.	Ammonia.
Acetic acid.	Lime-water.

[1] If time does not allow of this, filter a portion of the liquor.

Try the effect of each of these liquids on the blue litmus solution in the following way :—Fill one-third of a test-tube with water ; add 5 drops of the litmus solution by means of a pipette or piece of glass tubing ; then add 5 drops of each liquid, and stir. Carefully observe any change in colour by looking down the tube. State the effect you notice in a table thus :—

Liquid taken.	Effect on the Blue Litmus solution.
Sulphuric acid. Hydrochloric acid, etc.	Change to brick-red colour.

You have probably found that all the acid liquids turn the blue litmus solution **red**, while all the liquids possessing an alkaline instead of a sour taste leave the litmus solution **blue**, or render it still more blue. By this colour change, then, it seems that we can readily distinguish between an acid and an alkaline solution. But have no other liquids beside acid and alkaline solutions any colour effect upon litmus ? This question must be tested by experiment

Problem

Find whether blue litmus solution is affected by alcohol, benzene, a solution in water of common salt, washing-soda, or of any other substance you can think of. State your observations in a table.

Expt. 3. **To make Red and Blue Litmus Papers.**—Take a strong blue infusion of litmus, some very weak nitric acid, and some very weak caustic soda solution. Divide the blue litmus solution between two porcelain basins. To one portion add a few drops of the dilute acid till it is just reddened. To the other portion add one drop of the weak caustic soda solution so that it has a blue colour. Soak pieces of blotting-paper in each solution. Dry them as on p. 139. Then cut them into strips, and keep in a glass-stoppered bottle.

Problem

Try the colour-effects of acid, alkaline, and other liquids upon these red and blue litmus papers, by putting drops of the liquids upon strips of each kind by means of a glass rod. State what you observe in a table thus.

Liquid.	Effect on Red Litmus.	Effect on Blue Litmus.
Sulphuric acid. Hydrochloric acid, etc.	No visible change.	Red colour.

Expt. 4. **To try the Delicacy of the Litmus Test for an Acid.**—Add a cubic centimetre of strong sulphuric acid to a litre of water. Shake well. The acid is now diluted one thousand times. Put a drop on a blue litmus paper. It is reddened. Find whether the acid can still be detected when this weak solution is made 10 times and 100 times weaker still.

Expt. 5. **To prepare a Purple or Neutral Litmus Solution.**—Take some blue litmus solution in a porcelain basin, very weak acid solution, very weak alkali, and a pipette. Drop the dilute acid into the blue litmus solution, and stir with a glass rod until the colour is half-way between red and blue. Probably you will overstep the mark. If so, add the dilute alkali drop by drop very carefully. If the litmus changes in colour sharply from red to blue, weaken both the acid and alkaline solutions, and try again.

You will at last succeed in getting a **purple** or **neutral** litmus solution, intermediate in colour between red and blue.

To two portions of the purple or neutral litmus solution add acid and alkali respectively. One changes to red and the other to blue. Then, if we keep a stock of this one neutral solution, we can use it to detect both acids and alkalies, instead of using both red and blue solutions or papers.

Keep the neutral litmus solution in a glass-stoppered bottle.[1]

[1] If the solution loses its colour, the colour will be quickly restored on removing the stopper and shaking the solution with a little fresh air.

Problems

1. Prepare blue-purple and red-purple litmus solutions, intermediate in colour between blue and purple, and red and purple respectively.

2. Find whether lime-juice, soap, vinegar, washing-soda, bicarbonate of soda, common salt, nitre, borax, milk, sour milk, etc., have any effect on neutral litmus solution.

3. Find the effects of acids and alkalies upon solutions in alcohol of methyl orange and phenol phthalein; and upon infusions in water of red cabbage, logwood, and violet leaves. Fill one-third of a test-tube with water: add one drop of the solution, and then one drop of an acid or alkali by a pipette.

THE FORMATION OF SALTS

We have found that of the acids and alkalies known to us:—

1. An acid reddens blue litmus, and leaves red litmus unchanged.

2. An alkali turns red litmus blue, and leaves blue litmus unchanged.

What will be the effect of mixing an acid with an alkali in the presence of litmus?

Expt. 6. **To find the Result of adding Hydrochloric Acid to Caustic Soda until the Liquid is Neutral.**—Place about 10 c.c. of strong caustic soda solution in a basin. Add a *few* drops of litmus solution so that it is distinctly blue. Fill a pipette with dilute hydrochloric acid. Run the acid carefully drop by drop into the basin, stirring continually until the colour suddenly changes from blue to red. Then add some very dilute caustic soda solution until the liquid has a neutral or violet tint. Test also with litmus paper.

What is there now in the solution? Is it simply a mixture of the acid and the caustic soda? Suppose we find out what is left when the liquid is evaporated, for you will remember that hydrochloric acid, or spirit of salt as it used to be called, is volatile, so that we may expect to drive off the acid and have only the caustic soda left.

Evaporate the solution on wire gauze until the residue is dry. There will probably be some spurting, but this does not matter, as we do not need to weigh the residue in this experiment.

Do you observe the escape of any acid fumes? Does the solution still remain neutral as it evaporates?

Do you recognise the residue? Has it the appearance of caustic soda? Dissolve a little in water in a test-tube, and try whether the solution is alkaline or acid. Taste a very small portion. Do you not find it to be **common salt?**

Here, then, is an unexpected and important result. By the neutralisation of caustic soda with hydrochloric acid we do not obtain merely a mixture of the two with properties which are between those of the acid and the alkali. On the contrary, we find common salt to be produced, a substance which is very different in its properties from either of them. Is a similar result met with when other acids and alkalies are "neutralised?"

Problems

Find the products on neutralising—

1. 20 c.c. caustic potash with nitric acid.
2. 20 c.c. ammonia with hydrochloric acid.

Evaporate each solution, till the residue is dry, in a basin over a flame. Carefully examine and endeavour to identify the substances which you obtain by reference to appearance, taste, solvent action of water, and the action of heat in a dry test-tube.

Salts.—You have probably found after carrying out these problems that in each case a solid crystalline substance is obtained, which is neutral to litmus and has a salt-like taste. Other neutral mixtures of acids and alkalies might be made, and in each case you would obtain a neutral saline product. These substances are called **Salts.** The term "salt," then, as we shall use it in chemistry, applies not only to common sea-salt, but also to saltpetre and sal-ammoniac, and to many others, such as Glauber's salt and Epsom salts.

So far, then, as we can say at present a *salt is a substance, crystalline in appearance, saline in taste, neutral to litmus, which is obtained by the action of an acid upon an alkali.*

Old and New Names.—You have doubtless noticed that in the previous chapters we have sometimes given two names to the same substance—an old name, such as was used by the alchemists—and a new name adopted in more recent times. Fossils are records of the history of the progress of life on the

earth. Fossil names mark the progress of a science. For example "Spirit of salt," was the name given quite suitably by the alchemists to the volatile liquid obtained by distilling salt with oil of vitriol. The investigations of chemists during the last hundred years have however led to the abandonment of the name "spirit of salt" and the adoption of the name "hydrochloric acid." And because this latter name is the one that is probably to be found on the bottles in the laboratory, it will be more convenient to use it, although we cannot ourselves as yet attach any meaning to it, as we can do to "spirit of salt." Hence as a matter of convenience we shall adopt in many cases the more recent name, although as a matter of principle it would be better to retain the earlier name until we have found out for ourselves the reasons for employing what is really the more suggestive name.

The older names are still often used in commerce and in the household, while the more recent names find favour in the laboratory.

Names of Salts.—Salts are named after the acids from which they are formed. Thus—

Salts formed from	sulphuric acid	are called	sulphates.
,,	nitric acid	,,	nitrates.
,,	hydrochloric acid	,,	chlorides.
,,	acetic acid	,,	acetates.

The salt formed from caustic soda and hydrochloric acid is called *chloride of soda*. *Nitrate of potash* can be made from caustic potash and nitric acid. For reasons which will appear hereafter, chloride of soda is more often called *sodium chloride*, and nitrate of potash *potassium nitrate*.

Some of these substances may be recognised in taste and appearance as already familiar under other names.

It will be convenient to have a list of these names of salts and of some other substances at hand for reference :—

TABLE OF SYNONYMS

Old or Household Names	*New Names*
Oil of vitriol, or vitriolic acid	Sulphuric acid
Spirit of salt, or muriatic acid	Hydrochloric acid
Spirit of nitre, or aqua fortis	Nitric acid

Old or Household Names	*New Names*
Vinegar	Acetic acid
Washing-soda, mild alkali, or soda crystals	Sodium carbonate
Caustic soda, or caustic alkali	Sodium hydrate
Potash	Potassium carbonate
Caustic potash	Potassium hydrate
Sal-ammoniac	Ammonium chloride
Spirit of sal-ammoniac, hartshorn, or ammonia	Ammonium hydrate
Common salt	Sodium chloride
Nitre or saltpetre	Potassium nitrate
Chile saltpetre	Sodium nitrate
Epsom salts	Magnesium sulphate
Green vitriol	Ferrous sulphate
Blue vitriol	Copper sulphate
Chalk	Calcium carbonate
Quicklime	Calcium oxide
Slaked lime or lime water	Calcium hydrate
Fixed Air	Carbonic acid, carbon dioxide
Fire-Air	Oxygen
Spoilt Air	Nitrogen
Inflammable Air	Hydrogen
Nitrous Air	Nitric oxide
Volatile Sulphurous Acid	Sulphur dioxide

Problems

1. You are provided with a stick of caustic potash and with some strong nitric acid. Prepare a well crystallised sample of nitre.

2. Try to prepare dry and clean crystals of a salt—sulphate of potash—from a stick of caustic potash and some sulphuric acid.

3. Find whether the solutions **A**, **B**, **C**, are acid, alkaline, or neutral.

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CHAPTER XVIII

GRAPHIC REPRESENTATION[1]

Areas of Countries.—It is often difficult to read quickly figures which state the sizes of things, and the figures often give no clear idea to the mind. Thus it may be stated that—

Area of British Isles	=	121,115 sq. miles.
Area of Indian Empire	=	1,700,000 ,,
,, ,, under Native rule	=	750,000 ,,

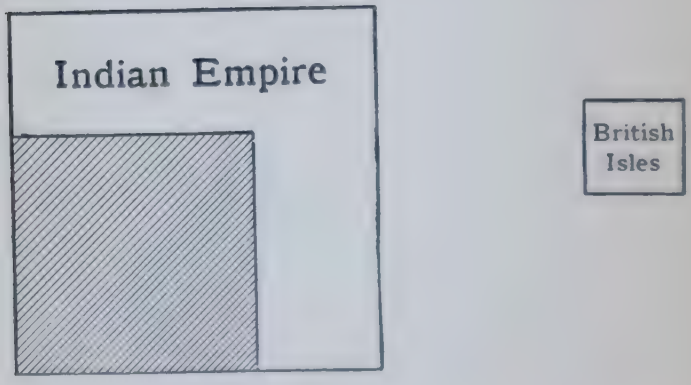


FIG. 90.

But a much clearer picture of these areas is obtained if a diagram is made showing squares drawn in proportion to these numbers as in Fig. 90.

[1] The authors call the attention of teachers to Prof. Perry's "Six Lectures at Jermyn Street in 1899" (Eyre and Spottiswoode, 6*d*.), from which they have obtained some valuable suggestions.

National Income.—Again, look at the following statement of the national income for the year 1900.

Excise	£33,100,000	Stamps	£7,825,000
Customs	26,252,000	Land Tax	755,000
Property and Income Tax	26,920,000	House Duty	1,720,000
Estate Duty	12,980,000	Crown Lands	500,000
Post Office	13,800,000	Suez Shares	830,000
Telegraph Service ...	3,450,000	Miscellaneous ...	2,243,000
			£130,384,000

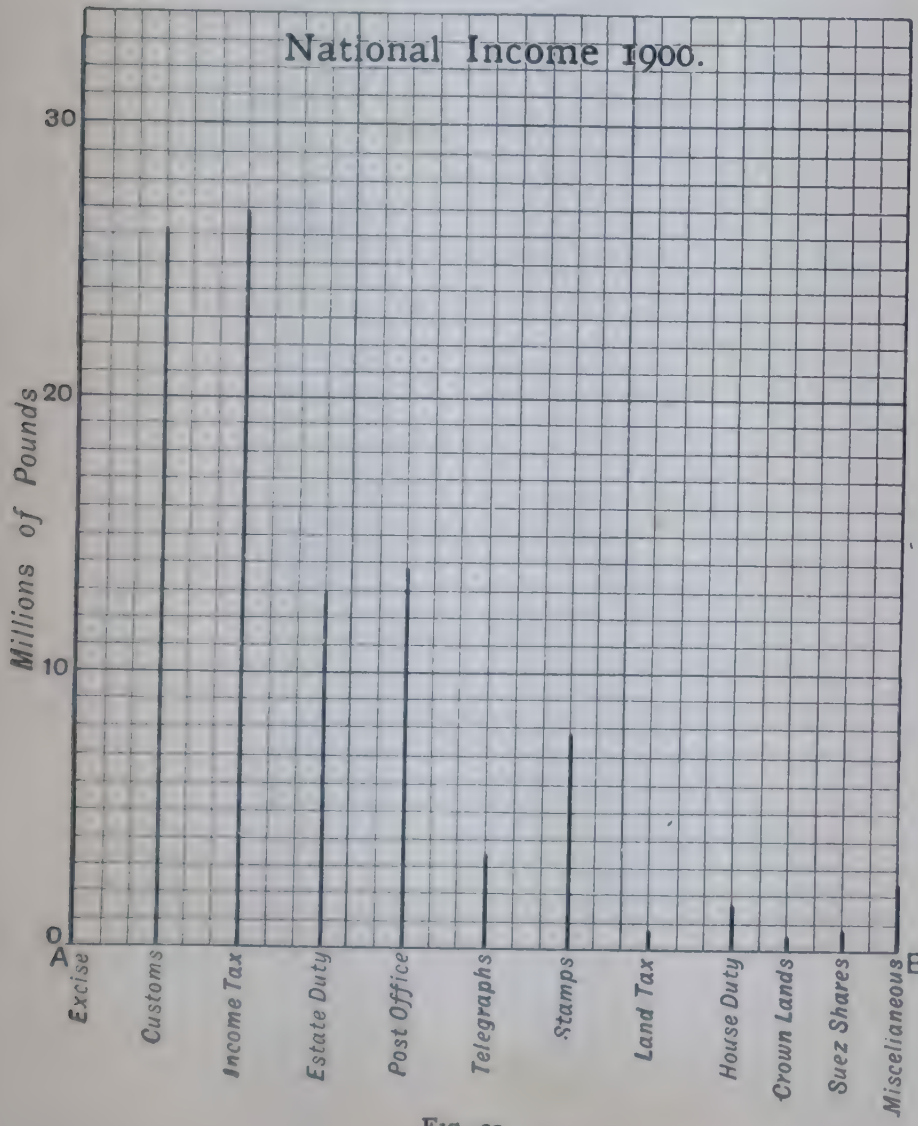


FIG. 91.

Large quantities like these are not easily grasped, but their relative sizes may be shown by drawing on paper lines proportional in length to each. The most convenient paper is what is called "**squared paper.**" This is covered with parallel horizontal and vertical lines at equal distances, and every tenth line is more distinct than the rest. On a sheet of squared paper draw a horizontal line, AB, along one of the thick lines. Then lines drawn at right angles to AB may be taken to represent millions of pounds. If, as is convenient, the paper is ruled in inches and tenths of an inch, a line one inch long may be taken to represent 10 millions of pounds. On this scale of a line one inch long representing 10 millions of pounds, proceed to draw with a sharp hard pencil straight lines at right angles to AB, to represent in order the above items of the nation's income (see Fig. 91). Test the accuracy of your diagram by reading from it the magnitudes of the several items, and comparing your readings with the above table.

Barometric Readings.—The following numbers show the height of the barometer at York, in inches, on each day of October, 1899, at 9 a.m.

October 1	29·45 inches.	October 17	30·34 inches.
2	29·60	18	30·39
3	29·78	19	30·45
4	29·79	20	30·44
5	30·14	21	30·42
6	30·19	22	30·27
7	30·20	23	30·06
8	30·39	24	29·95
9	30·27	25	29·89
10	30·18	26	29·84
11	30·01	27	29·76
12	29·87	28	29·77
13	29·71	29	29·70
14	30·16	30	29·75
15	30·23	31	29·84
16	30·29		

These barometric heights might be represented by a row of 31 straight lines, as in Fig. 91, all standing in order upon one horizontal line, but it will be sufficient to mark the *tops* only of these lines by points. Moreover, since each height exceeds

29 inches, we need only represent the heights *above* 29 inches, and so make the diagram more compact.

Begin, therefore, by drawing two lines OX and OY at right angles to one another (see Fig. 92), along two of the thick lines upon squared inch paper. (These are often called the *axes*). Mark every tenth line along OX, 10, 20, 30, and along OY mark O as 29, and the next thick line 30, and the next 31. Proceed to mark points on the paper with a sharp-pointed hard pencil to represent the above barometric heights.

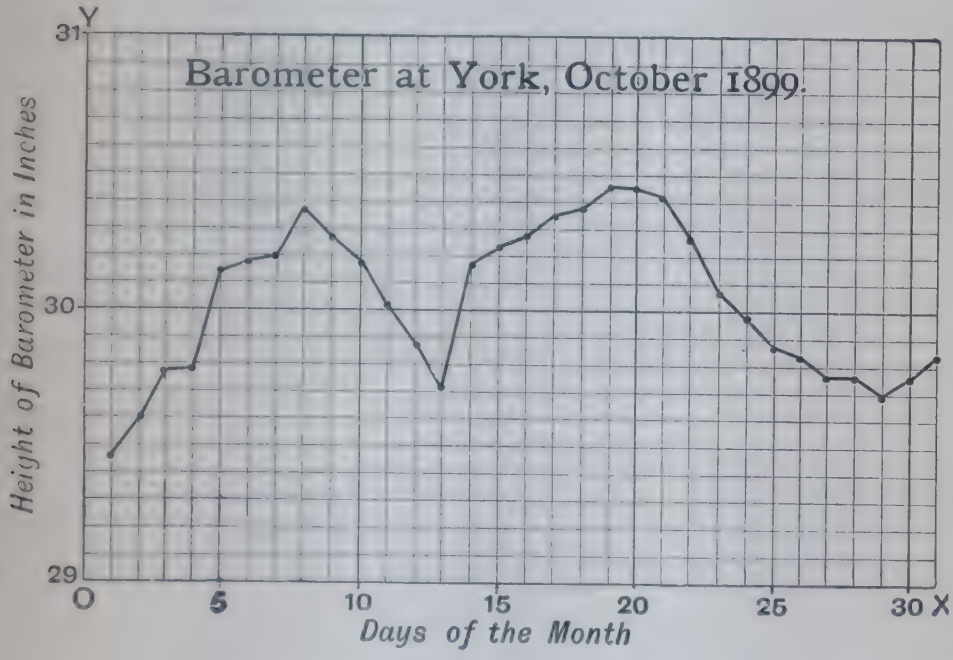


FIG. 92.

Further, since the barometer changes continuously from one height to the next, we may indicate this by joining the points by fine, short, *straight* lines, as in Fig. 92, for according to the data we do not know how the barometric height may have varied during each twenty-four hours. The irregular line so obtained is called a **barometric curve.**

Reference to a self-recording barometer would probably show that the height did not change between the successive readings quite regularly, as our curve implies. Sometimes the barometer is read more often than once a day. Thus at the office of the

Daily Chronicle in London it is read four times a day ; the barometric curve consequently takes a more wavy appearance. This is shown by the thick wavy line in Fig. 93, which shows a *Daily Chronicle* weather chart for four days.

Observe that there was no reason to join the tops of the straight lines in the Revenue Diagram (Fig. 91), because the lines in it represent *different* things, not one thing which varies continuously in magnitude. So also if we were to plot the daily

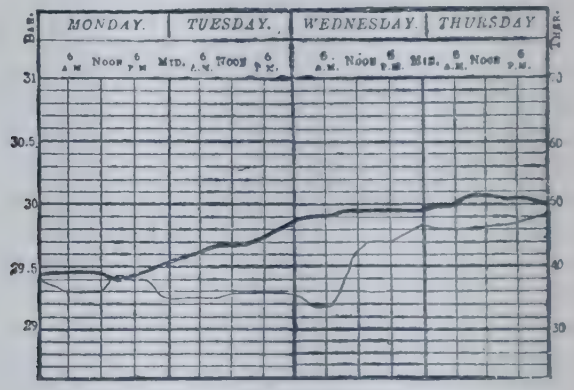


FIG 93.[1]

rainfall we should not attempt to draw a curve connecting the measurements, since the quantity of rain changes abruptly, not continuously, from one day to another.

EXERCISES

In the following exercises plot the tabulated quantities on squared paper. In all cases state on the paper what it is you have represented, and also name the different scales. In choosing the scales avoid those which would make the diagrams either too large, too long or broad, or too crowded.

[1] From Gregory and Simmons's *Elementary Physics and Chemistry*, Part I., p. 141.

1. *Annual Expenditure of Families.*

	A. Average of Twenty Artisan Families.	B. Average of Twelve Middle Class Families.
Rent	£15	£46·4
Food	48·5	101
Clothing	8·8	51·3
Fire and Light	4·8	17·2
Doctor	1·5	9·3
Education	—	16·3
Travelling and Holidays ...	13·6	41·6
Charity	—	24·3
Insurance	—	29·1

2. *Food Expenditure in a Boarding School*, 1899.

Flour	£234·6	Dried Fruits	£24·0
Meat and Fish	1011·6	Vegetables and Fresh Fruit	314·3
Milk	594·3	Coffee and Cocoa	24·4
Butter	228·3	Sugar and Syrup	106·0
Cheese	11·0	Rice	10·8
Tea	49·9	Sundries	174·0

3. *Expenditure of the British and Foreign Bible Society.*

(Founded 1804.)

1804	£692	1854	£149,000
1814	81,000	1864	169,700
1824	94,000	1874	217,400
1834	84,200	1884	231,000
1844	85,800	1894	215,000

4. *Punishment Lines given in a School in successive weeks of a Term.*

Week 1	1050 lines.	Week 8	2750 lines.
2	2190	9	2695
3	2210	10	3535
4	2000	11	2900
5	2170	12	4360
6	1845	13	4490
7	3280	14	4230

5. *Date of Appearance of Wild Flowers.*[1] (Average of 10 years.)

The numbers give the date counting from January 1; *e.g.*, the numbers 1–31 indicate dates in January.

Hazel 32	Marsh Marigold 77	Purple Orchis ... 124
Coltsfoot 52	Cowslip 81	Hawthorn... ... 135
Primrose 54	Cuckoo-flower.. 95	Red Clover ... 137
Lesser Celandine 60	Bluebell 113	Forget-me-not ... 148
Sweet Violet ... 71	Crab 116	White Clover ... 155
Anemone 76	Upright Buttercup 116	Wild Rose ... 170

6. *Barometric Readings at York, January* 1884.

Jan 1 ..... 30·44 ins.	Jan. 11 ..... 29·81 ins.	Jan. 21 ..... 30·30 ins.
2 ..... 30·16	12 ..... 30·32	22 ..... 29·98
3 ..... 29·89	13 ..... 30·39	23 ..... 29·63
4 ..... 30·09	14 ..... 30·30	24 ..... 29·62
5 ..... 29·98	15 ..... 30·49	25 ..... 29·41
6 ..... 29·61	16 ..... 30·60	26 ..... 29·09
7 ..... 29·83	17 ..... 30·54	27 ..... 28·67
8 ..... 30·04	18 ..... 30·54	28 ..... 29·25
9 ..... 30·08	19 ..... 30·48	29 ..... 29·68
10 ..... 30·17	20 ..... 30·28	30 ..... 29·54
		31 ..... 29·78

7. *Barometric Readings at York, December* 1886.

Dec. 1 ..... 29·77 ins.	Dec. 11 ..... 29·31 ins.	Dec. 21 ..... 30·35 ins.
2 ..... 29·86	12 ..... 29·08	22 ..... 29·80
3 ..... 30·02	13 ..... 29·69	23 ..... 29·62
4 ..... 29·64	14 ..... 29·59	24 ..... 29·47
5 ..... 29·96	15 ..... 29·18	25 ..... 29·78
6 ..... 29·62	16 ..... 29·33	26 ..... 29·81
7 ..... 29·21	17 ..... 29·64	27 ..... 29·66
8 ..... 28·34	18 ..... 29·58	28 ..... 29·69
9 ..... 28·18	19 ..... 29·74	29 ..... 29·78
10 ..... 29·08	20 ..... 30·04	30 ..... 30·30
		31 ..... 30·56

8. *Maximum and Minimum Temperatures.*

Observed at Manchester, during parts of September and October, 1805. Plot on the same paper.

	Maximum.	*Minimum.*
Sept. 20	64°·7 F.	44°·4 F.
21	65°·2	43°·0
22	70°·3	44°·0
23	75°·8	43°·7
24	81°·2	56°·4
25	79°·4	61°·6
26	78°·5	54°·5

[1] At Ackworth, Yorkshire. From *The Natural History Journal*, 1888, p. 116.

		Maximum.	*Minimum.*
Sept.	27	81°·5	58°·1
	28	83°·9	55°·3
	29	81°·4	52°·5
	30	76°·9	52°·2
Oct.	1	71°·9	52°·4
	2	52°·0	47°·3
	3	58°·8	36°·1

9. *Sunshine and Rain.*

Observed at Manchester, during parts of September and October, 1895.

		Sunshine. Hrs.	Mins.	*Rain.* Ins.
Sept.	20	4	38	—
	21	7	14	—
	22	8	48	—
	23	5	56	—
	24	6	15	0·041
	25	4	14	0·017
	26	4	30	0·032
	27	7	40	—
	28	7	56	—
	29	8	32	—
	30	6	32	—
Oct.	1	1	24	—
	2	1	8	0·792
	3	0	43	0·226

10. *Average Heights and Weights of Children.*[1]

Age last birthday.	Boys. Height.	Weight.	Girls. Height.	Weight.
5 years	41·74 ins.	41·20 lbs.	41·47 ins.	39·82 lbs.
6	44·10	45·14	43·66	43·81
7	46·21	49·47	45·94	48·02
8	48·16	54·43	48·07	52·93
9	50·09	59·97	49·61	57·52
10	52·21	66·62	51·78	64·09
11	54·01	72·39	53·79	70·26
12	55·78	79·82	57·16	81·35
13	58·17	88·26	58·75	91·18
14	61·08	99·26	60·32	100·32
15	62·96	110·84	61·39	108·42
16	65·58	123·67	61·72	112·97
17	66·29	128·72	61·99	115·84
18	66·76	132·71	62·01	115·80

[1] From Annual Report of State Board of Health, Massachusetts, 1877.

Pressure of Water Vapour.— On page 98 a table was given of the pressure of water vapour measured in millimetres

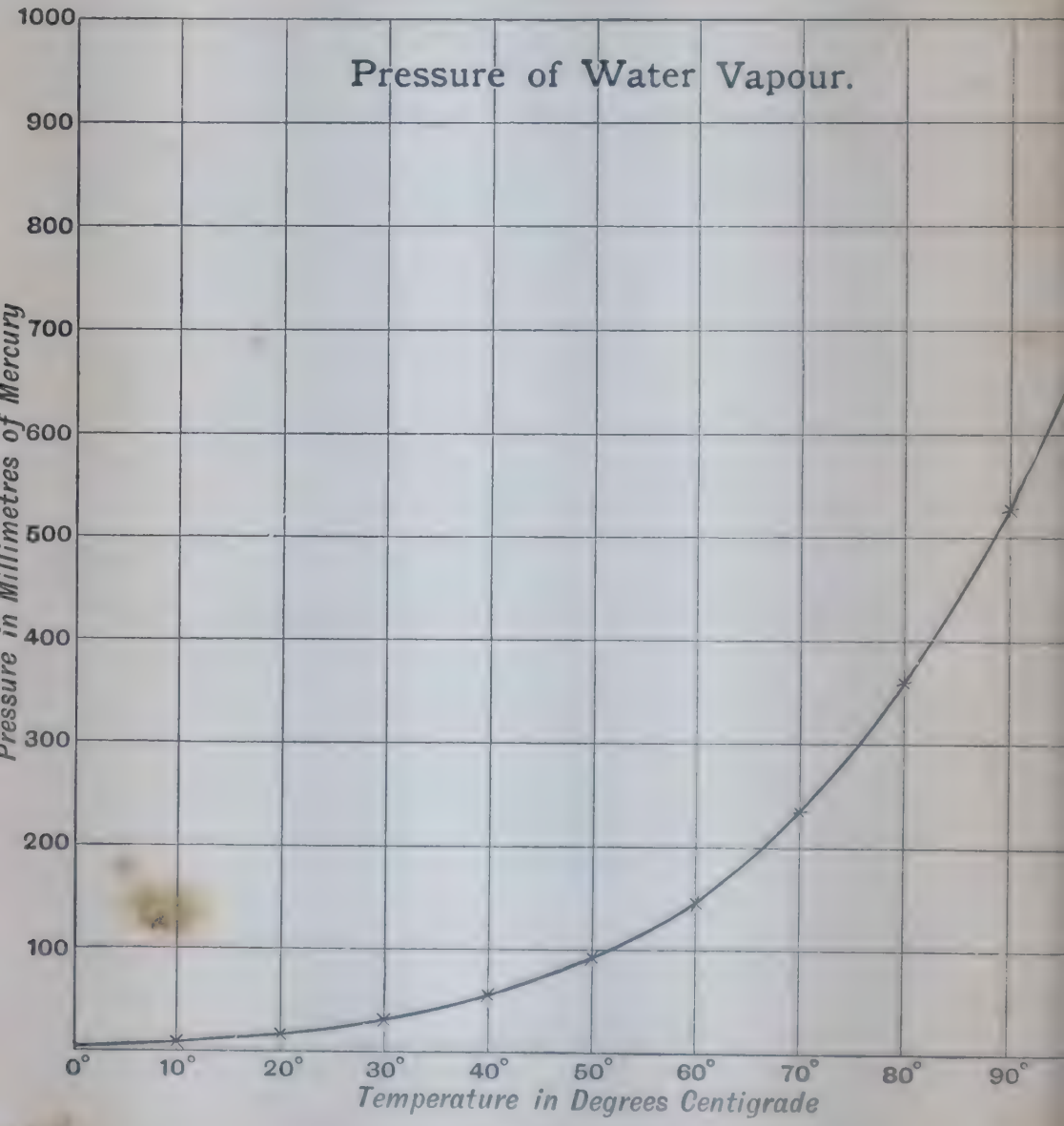


FIG. 94.

of mercury at various temperatures. This may be repeated here and somewhat extended.

Temperature	Pressure
0° C.	4·+ mm.
10° ,,	9·+ ,,
20° ,,	17·+ ,,
30° ,,	31·+ ,,
40° ,,	54·+ ,,
50° ,,	91·+ ,,
60° C.	148·+ mm.
70° ,,	233·+ ,,
80° ,,	354·+ ,,
90° ,,	525·+ ,,
100° ,,	760·+ ,,
105° ,,	906·+ ,,
110° ,,	1075·+ ,,

Proceed to mark these measurements by means of small crosses upon squared millimetre paper, marking degrees Centigrade off along a horizontal axis, and mm. of mercury up from it along a vertical axis (see Fig. 94). Observe that a curved line may be drawn sweeping through the points marked by these crosses. Endeavour to draw the curve.[1] Any fault in the curve can be more readily detected if the paper is held in such a way that the eye can look *along* it, close to the paper. Observe (see Fig. 94)—

(1) That the *slope* of the curve becomes steeper and steeper as the temperature rises: this shows very clearly that the pressure of water vapour increases much faster as the temperature rises.

(2) That the curve enables us to find the pressure of water vapour at temperatures lying between those at which actual measurements have been made: *e.g.* the curve shows that the pressure of water vapour at 85° is measured by 430 mm. of mercury. Values obtained in this way are said to be obtained by **interpolation.** Thus the careful plotting of a few accurate measurements enables us to obtain many other values

Solubilities of Solids.—It has been found in Chapter XVII. that the solubility of a solid in water usually changes with change of temperature. This is seen in the following statement of the quantities of nitre, salt, and chlorate of potash, which can be dissolved in 100 grams of water at certain different temperatures:—

[1] *Note on Drawing a Curve through a Number of Points.*

(i.) Put needles in at the points marked: bend a strip of whalebone or a straight edge round them, and then rule the line with a pencil.

(ii.) Draw the curve freehand, keeping the wrist on the *inside* of the curve whilst drawing, so that the natural motion of the hand round the wrist as centre may help in forming the curve. Sketch the line with a pencil.

	Nitre.	Salt.	Chlorate of Potash.
0° C.	13 gm.	35·5 gm.	3 gm.
10°	21	35·8	4
20°	31	36·1	6
30°	45	36·4	9
40°	64	36·6	13
50°	86	36·9	18
55°	100		
60°		37·2	24
70°		37·5	32
80°		37·8	40
90°		38·1	49
100°		38·4	60

The meaning of these numbers, each one the result of a careful experiment, can be best seen if they are plotted on squared paper.

Along a horizontal axis on a sheet of squared millimetre paper, mark off *temperatures* from 0° to 100° C., and mark off *solubilities* (*i.e.* grams of solid dissolved in 100 grams of water), from 0 to 100 grams up a vertical axis. Proceed to mark the measurements of the solubility of *nitre* by crosses upon the paper. Observe that a curved line might be drawn sweeping through the points so marked. Endeavour to draw the curve. This line is called the **solubility curve of nitre.** Similarly proceed to draw on the same sheet of paper the solubility curves for *salt* and *chlorate of potash* (see Fig. 95).

A solubility diagram is not only a record of experiments: it also supplies the key to many problems. This will be understood after working through the following examples:—

Examples XIII

[*To be answered by reference to the diagram* you *have made.*]

1. How many grams of nitre will dissolve in 100 grams of water at 4°, at 35°, at 45°?

2. At what temperatures will 100 grams of water be just saturated by 5 grams of potassium chlorate, by 19 grams, by 55 grams?

3. How much more chlorate of potash will dissolve in 100 grams of water (i.) at 60° than at 15°, (ii.) at 80° than at 55°?

4. At what temperature are salt and chlorate of potash equally soluble in water?

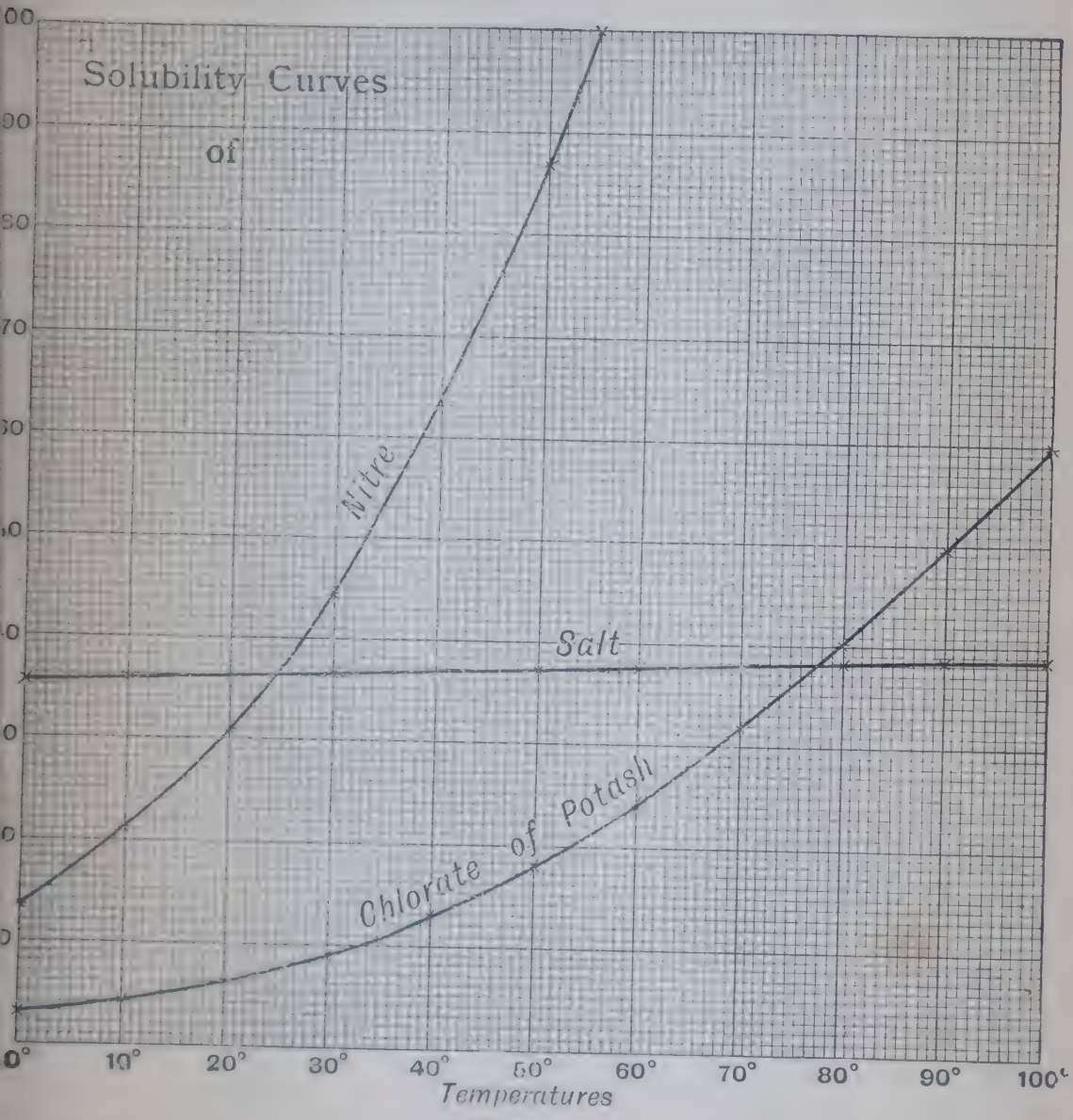


FIG. 95.

5. Which of the three substances is the most soluble (i.) at 50°, (ii.) at 10°?

6. By how many grams per degree does the solubility of chlorate of potash increase between 80° and 100°?

7. At what temperature will 100 grams of water just dissolve (i.) 25 grams of nitre, (ii.) 25 grams of potassium chlorate?

8. How much water will just dissolve (i.) 90 grams of nitre at 30°, (ii.) 20 grams of chlorate of potash at 80°?

9. How many grams of nitre will crystallise out when a solution of nitre in 100 grams of water, saturated at 35°, cools to 10°?

10. How could I make a solution containing 25 grams of chlorate of potash per 100 grams of water, using a thermometer but without a balance?

Volume of 1 gram of Water at Different Temperatures.—It is a fact of far-reaching importance that water has

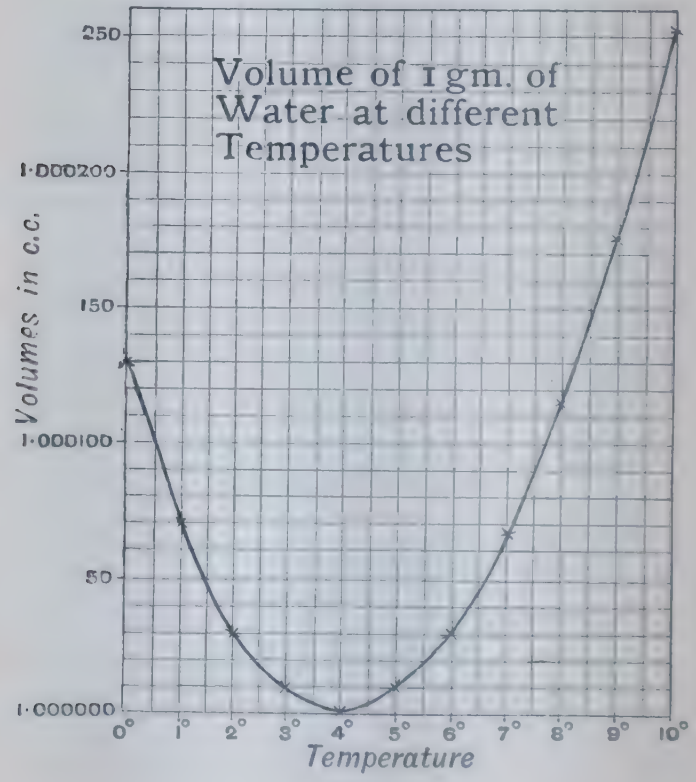


FIG. 96.

at 4° C. a greater density than at 0° C. This behaviour of water, as it approaches its freezing-point, is an exception to the almost universal fact that liquids contract on cooling. Water at 4° C. if heated expands, as we should expect. Now 1 gram is by definition the weight of 1 c.c. of water at 4° C., and therefore

1 gram of water below or above 4° C. occupies more than 1 c.c. The following table shows the volume of 1 gram of water at various temperatures :—

0° C.	1·000129 c.c.	6° C.	1·000030 c.c.
1°	1·000072	7°	1·000067
2°	1·000031	8°	1·000114
3°	1·000009	9°	1·000176
4°	1·000000	10°	1·000253
5°	1·000010		

These volumes differ very slightly in actual magnitude, only in the last three places of decimals ; but nevertheless, on plotting them, such scales may be chosen that the diagram shall be of a reasonable size, and so that the differences between the volumes shall be evident. It will be necessary to magnify the scale representing the volumes. Use squared inch paper ; mark temperatures along a horizontal axis, and mark volumes along a vertical axis, letting each 0·1 inch represent 0·000010 c.c Draw a curve through the points (see Fig. 96).

Times of Sunrise and Sunset.—An almanac predicts the following times of sunrise and sunset for a part of September :—

	A. *Sunrise.*	B. *Sunset.*
Sept. 16	5.38 a.m.	6.12 p.m.
17	5.40	6.9
18	5.41	6.7
19	5.43	6.5
20	5.44	6.2
21	5.46	6.0
22	5.48	5·58
23	5.49	5·55
24	5.51	5·53
25	5.53	5·51
26	5.54	5.48
27	5 56	5.46
28	5·57	5·44
29	5·59	5·42
30	6.1	5·39

Plot the predicted times of *sunrise* on successive days upon squared paper, marking days of September, starting with the

15th along a horizontal axis, and times starting with 5.30 vertically up from it. (Express the minutes as decimals of an hour.) Join the points marked, and observe that the joining line is not a regular curve (see Fig. 97, AA). Why is this? The times tabulated are not quite accurate; they are only stated to the nearest whole minute. And, indeed, it will frequently be

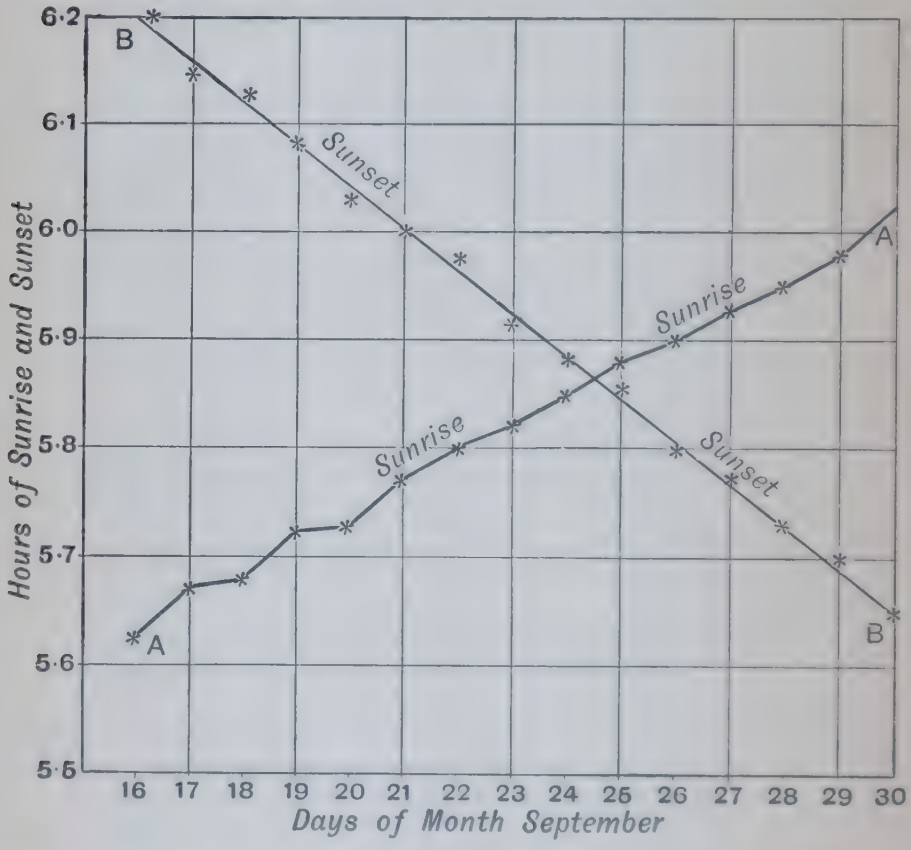


FIG. 97.

found that a series of *not quite accurate* statements will give an irregular curve such as this.

So, also, laboratory experiments involving measurements are certain to contain some errors, and if the observations are plotted, the joining line will be somewhat irregular.

In such cases, if we are confident that the points ought to be connected by a smooth line, we may *draw evenly among the points a smooth line*. Proceed now to plot the predicted times of *sunset*: observe that a straight line may be drawn passing

evenly *among* the points (see Fig. 97, BB), *i.e.* so that there are as many points on one side of the line as on the other. This line represents the most probable connection between the times of sunset, and even *enables us to see which of the tabulated times are incorrect*, either too early or too late.

Price Lists.—The price of any article frequently varies from time to time. A record of the price, which can be read at a glance, is afforded if the prices are plotted. Thus Fig. 98 is a record of the *price of india-rubber* from 1877 to 1898. It shows

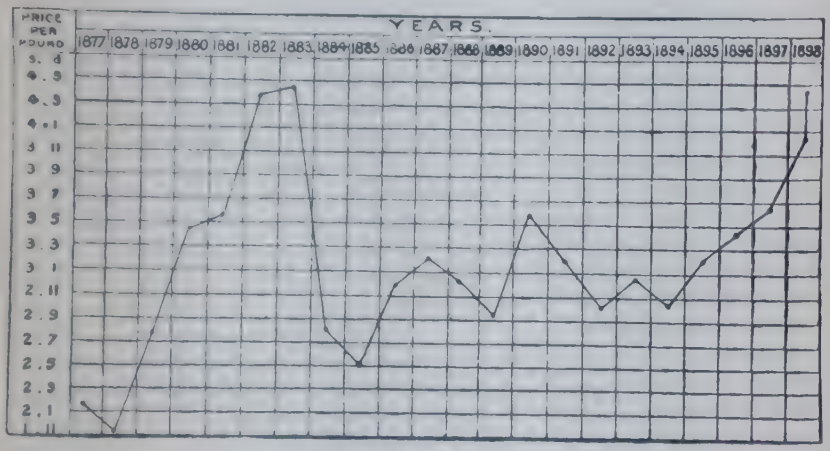


Fig. 98.[1]

that in 1877 the price per lb. was about 2*s.* 2*d.*, and in 1878 1*s.* 11*d.*, the lowest price in the series of years. In what years was a maximum price reached? 1883 and 1898. Between what two years was there the greatest change in the price? 1883–84.

Again, the price of any one kind of manufactured article varies with its size. Suppose a manufacturer of *lantern screens* wishes to publish a price list of various sizes. Suppose he has made screens 6, 9, 12, and 15 feet square, and arranged their prices thus :—

a screen 6 feet square	...	...	...	£1	5	0	=	£1·25
,, 9 ,,	,,	...	...	...	1	17	6	=	1·875
,, 12 ,,	,,	...	...	...	3	10	0	=	3·50
,, 15 ,,	,,	...	...	...	6	0	0	=	6·00

[1] Quoted by Gregory and Simmons, from the *Kew Bulletin*, in *Elementary Physics and Chemistry*.

Let him plot these sizes and prices on squared paper as in Fig. 99, and draw a curve through the points.

Any point on the curve then shows the size and the probable price of a lantern screen. The manufacturer can now read off

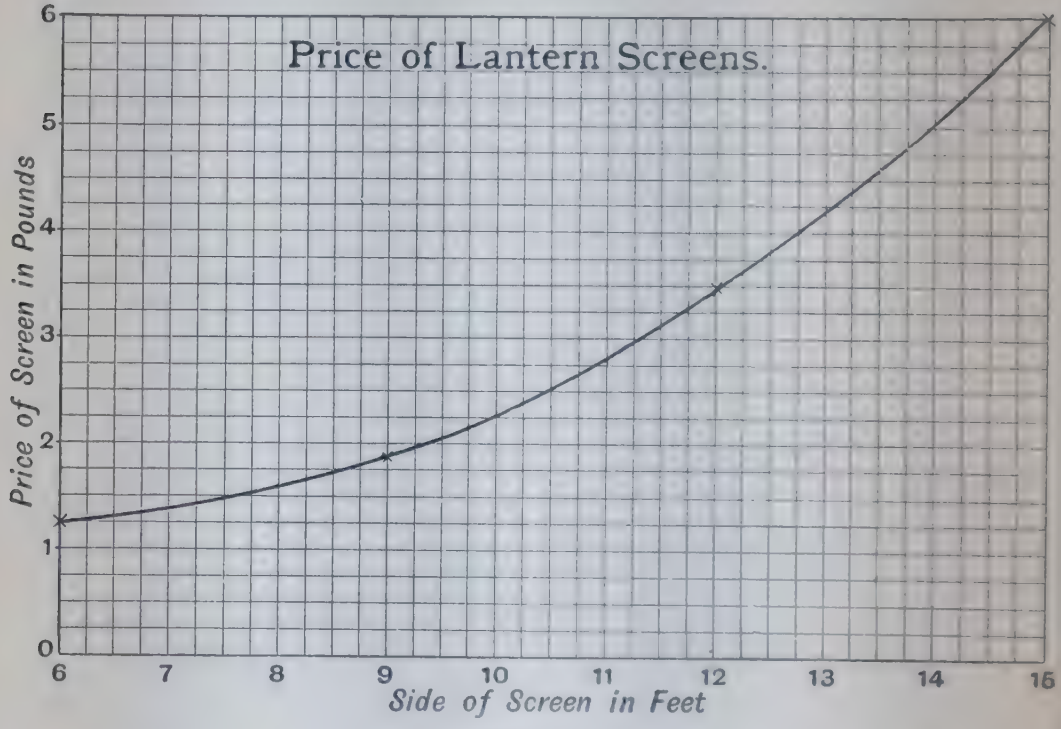


FIG. 99.

the probable prices of other sizes of screen ; for instance, he finds—

a screen	7½	feet	square	should	be	priced	...	...	£1	9 0
,,	10	,,	,,	,,	,,		...	...	2	6 0
,,	14	,,	,,	,,	,,		...	...	5	0 0

This calculation of prices of screens in size lying *between* those of which the prices are known is another illustration of the use of interpolation.

Population. The following numbers give the population of England and Wales at intervals of ten years during the 19th century :—

1801	... ...	8,893,000	
1811	... ...	10,164,000	
1821	... ...	12,000,000	
1831	... ..	13,897,000	

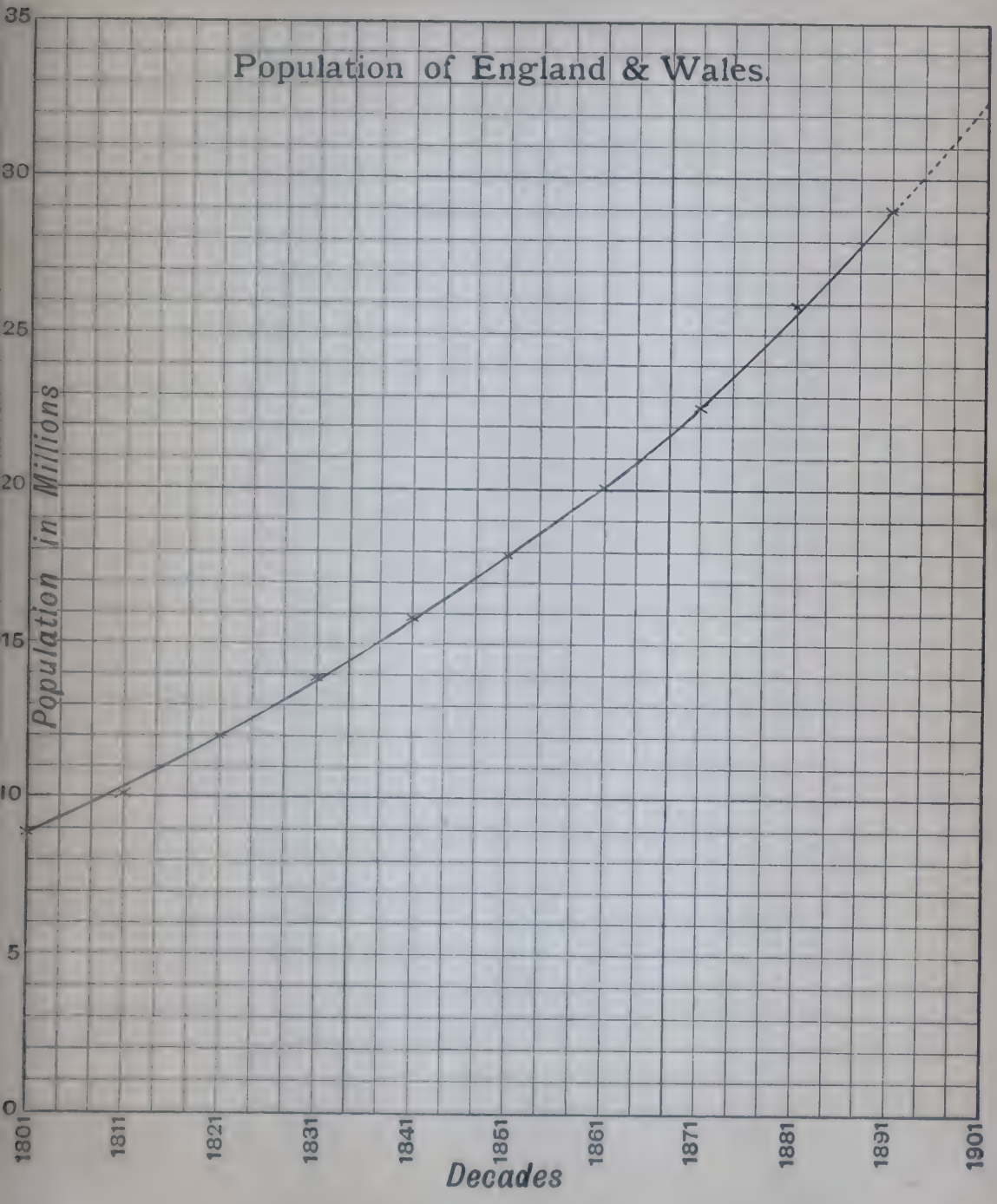

Population of England & Wales.
Population in Millions
35
30
25
20
15
10
5
0
1801
1811
1821
1831
1841
1851
1861
1871
1881
1891
1901
Decades

FIG. 100.

1841	15,914,000	1871	22,712,000
1851	17,928,000	1881	25,974,000
1861	20,066,000	1891	29,003,000

Plot these numbers on squared paper, and draw a curve lying evenly *among* the points marked, as in Fig. 100.

From the curve we can state what was the probable population of England and Wales in any particular year between 1801 and 1891. But also, by continuing the curve beyond the last plotted point corresponding to 1891, we can find out what it was likely to be in 1901, ten years after the last census in the Table. We see that in 1901 the population of England and Wales was likely to be about $32\frac{1}{2}$ millions.[1] Observe also that the *slope* of the curve becomes steeper and steeper as the years advance; this means, of course, that the *rate of increase* of the population becomes greater and greater.

Paper for curve plotting is manufactured in squares of $\frac{1}{10}$ inch side, with every tenth line thickened; also in square millimetres and centimetres; also in larger squares as exercise books by the educational stationers.

Paper ruled on the English scales, the square inch and its fractions, in great variety, of the best, is made by Messrs. Waterlow and Sons, Limited, 85 London Wall, E.C. Good ten-to-the-inch paper may be obtained from The Scientific Publishing Co., 53 New Bailey Street, Manchester.

Good millimetric rulings may be obtained through Messrs. Williams and Norgate, foreign booksellers, Henrietta Street, Covent Garden.

Some of the scientific instrument dealers stock useful qualities, *e.g.*, Harris of Birmingham, and Reynolds and Branson of Leeds. The paper sold by the educational stationery shops is less accurate. John Bellows, Gloucester, can do good work, to order.

Large sheets of "design paper," not quite so finely ruled, but in great variety of squares and oblongs, are used for lace designing in Nottingham, and may be obtained through stationers in that town, *e.g.*, Messrs. Mounteney.

[1] According to the census taken on April 1, 1901, the population of England and Wales is now (1901) 32,526,075.

EXERCISES: CURVE-PLOTTING

In the following exercises plot curves to represent the tabulated quantities. Those scales should be avoided which are obviously inconvenient. Scales should be so chosen that the plotted curve is not crowded in one corner of the paper. No exercise is completed until the scales and the names of the plotted quantities are clearly indicated on the paper.

1. *Grams of Solid dissolved in* 100 *grams of water.*

Temperature.	Barium Nitrate.	Potash Alum.	Sodium Sulphate.
0° C.	5·2	3·9	5·0
10°	7·0	9·5	9·0
20°	9·2	15·1	19·4
30°	11·6	22·0	40·0
40°	14·2	30·9	48·8
50°	17·1	44·1	46·7
60°	20·3	66·6	45·3
70°	23·6	90·7	44·4
80°	27·0	134·5	43·7
90°	30·6	209·3	43·1
100°	32·2	357·5	42·5

2. *Pressure of Water Vapour above* 100° C.

Temperature.	Pressure.	Temperature.	Pressure.
100° C.	1 atmosphere.	165°·3 C.	7 atmospheres.
112°·2	1½ ,,	198°·8	15 ,,
120°·6	2 ,,	201°·9	16 ,,
133°·9	3 ,,	204°·9	17 ,,
144°	4 ,,	207°·7	18 ,,
152°·3	5 ,,	210°·4	19 ,,
156°·2	6 ,,	213°·0	20 ,,

3. *Density of Water between* 0° *and* 20° C.

0° C.	0·99988	12° C.	0·99955
2°	0·99997	14°	0·99930
4°	1·00000	16°	0·99900
6°	0·99997	18°	0·99840
8°	0·99988	20°	0·99807
10°	0·99974		

4. *Price List of Bottles*, white glass, English, stoppered.

Capacity : oz.	½	1	2	3	4	6	8	10	12	16	20
Price per doz.	4/-	4/6	5/6	6/-	7/-	8/-	8/6	10/6	12/-	14/-	15/-

Does it appear that the price list has been correctly made out?

5. *Price List of Basins*, porcelain, Berlin, for evaporation.

Diameter : inches	$2\frac{3}{4}$	$3\frac{1}{4}$	$3\frac{1}{2}$	$3\frac{3}{4}$	4	$4\frac{1}{4}$	$4\frac{3}{4}$
Price, each... ...	-/4	-/6	-/8	-/10	1/-	1/2	1/6

6. *Price List of Rubber Corks*, solid.

Diam. small end, in inches	$\frac{3}{8}$	$\frac{1}{2}$	$\frac{5}{8}$	$\frac{3}{4}$	$\frac{7}{8}$	1	$1\frac{1}{8}$	$1\frac{1}{4}$
Price per doz.	1/6	1/9	2/4	3/-	4/3	4/9	6/-	7/9
Diam. small end, in inches	$1\frac{3}{8}$	$1\frac{1}{2}$	$1\frac{5}{8}$	$1\frac{3}{4}$	2	$2\frac{1}{4}$	$2\frac{1}{2}$	$2\frac{3}{8}$
Price per doz.	9/-	10/-	11/3	12/6	16/-	22/-	33/-	40/

7. *Price List of Filter Papers*, Rhenish.

Diameter in cm.	4·25	5·5	7	9	11	12·5	15
Price per 1000	2/9	2/9	3/-	4/-	4/9	5/-	7/3
Diameter in cm.	18·5	24	27	32	38·5		
Price per 1000	8/9	15/9	22/-	26/10	31/6.		

8. *Population of Scotland.*

1811	1,806,000	1861	3,062,000
1821	2,091,000	1871	3,360,000
1831	2,364,000	1881	3,736,000
1841	2,620,000	1891	4,026,000
1851	2,889,000		

Estimate the population of Scotland in 1901.

9. *Premiums for Life Insurance.*—To insure the payment of £100 at death, the Government requires the following yearly premiums to be paid throughout life. (Change the shillings and pence into decimals of £1.)

Age of Insurer.	Premiums.	Age of Insurer.	Premiums.
15	1 8 6	40	2 18 0
20	1 13 0	45	3 9 6
25	1 17 6	50	4 4 0
30	2 3 0	55	5 4 0
35	2 9 6	60	6 10 6

What premium should be paid if the insurance is made at the age of 10 years?

10. *Conversion of Fahrenheit and Centigrade Temperatures.*

32° F. = 0° C.	140° F. = 60° C.
50° = 10°	158° = 70°
68° = 20°	176° = 80°
86° = 30°	194° = 90°
104° = 40°	212° = 100°
122° = 50°	

By interpolation find

(*a*) What C. temperatures correspond to 41° F. and 77° F.
(*b*) ,, F. ,, ,, 15° C. ,, 95° C.

11. *Logarithms of Numbers.*

Logarithm of 310 = 2·49136
,, ,, 311 = 2·49276
,, ,, 312 = 2·49415
,, ,, 313 = 2·49554
,, ,, 314 = 2·49693
,, ,, 315 = 2·49831

By interpolation find the logarithms (to 4 places of decimals) of 310·4, 311·8, 312·55, 314·12, and 314·86.

12. *Logarithms of Numbers.*

Number.	Logarithm.	Number.	Logarithm.
1	0·000	6	0·778
2	0·301	7	0·845
3	0·477	8	0·903
4	0·602	9	0·954
5	0·698	10	1·000

By interpolation find the logarithms (to 2 places of decimals) of 1·5, 2·3, 5·8, 8·1, 9·26.

13. *Times for lighting Cycle Lamps.*

January 21	5.28 P.M.	July 21	9.2 P.M.
February 21 ...	6.24	August 21	8.10
March 21	7.12	September 21 ..	7.1
April 21	8.4	October 21... ...	5.5
May 21	8.50	November 21 ...	5.3
June 21	9.19	December 21 ...	4.51

14. If a smooth curve can be drawn among a series of plotted points, it is probable that there is a relation between the two sets of quantities plotted. Find whether there is a relation between the *Distances of the Planets from the Sun, and their Times of Revolution round the Sun*

	Distance in Millions of Miles.	Times of Revolution in Days.
Mercury	35	87
Venus	66	224
Earth	92	365
Mars	141	686
Jupiter	480	4,332
Saturn	881	10,759
Uranus	1,771	30,688
Neptune	2,775	60,181

15. Find whether there is any connection between the quantities tabulated below :—

	Specific Heat.	"Atomic Weight."
Lithium	0·94	7
Sodium	0·29	23
Aluminium	0·20	27
Potassium	0·166	39
Iron	0·112	56
Zinc	0·093	65
Silver	0·056	108
Tin	0·054	118
Iodine	0·054	127
Platinum	0·032	195
Mercury	0·032	200

16. *The Freezing-Point of Carbolic Acid.*—Fill one-quarter of a narrow test-tube with pure crystallised carbolic acid. Plunge the bulb of a thermometer into it and fix the tube nearly horizontally in a clamp. Warm the tube very gently until all the carbolic acid is melted, but do not let the temperature rise above 50° C. Then allow the tube to cool, and read the temperature accurately every half-minute, until it falls to nearly 30° C. Between the readings stir gently with the thermometer, but take care not to take the bulb out of the carbolic acid. Plot your observations, marking temperatures along the vertical axis. What temperature does the curve show to be the freezing-point of carbolic acid?[1]

17. *The Freezing-Point of Sulphur.*—Proceed as in No. 16, but take readings every 15 seconds, as the temperature falls from 130° to 100° C.

18. *The Freezing-Point of Paraffin Wax.*—Proceed as in No. 16, but take readings every 15 seconds, as the temperature falls from 60° to 40° C.

Note.—Interesting additional exercises may be obtained from Whitaker's Almanac, Chambers's Mathematical Tables, the Nautical Almanac, and from the daily newspapers.

[1] The *freezing-point* of a substance is the temperature at which a substance remains as it is changing from the solid to the liquid state.

EXAMINATION PAPERS

The following papers have all been actually set in schools as terminal examinations upon the subject-matter of this volume. Not more than two hours were allowed in each case. The *average* age of the Forms examined was not less than $13\frac{1}{2}$ years nor more than 15 years

(*a*)

1. What roughly represents (*a*) length of 1 decimetre?
(*b*) volume of 1 cubic centimetre?
How many grams are there in 1 ounce?

2. Add, expressing the answer in grams and decimal of a gram:—
1 Kgm. + 20 mg. + 3 C. + 5 D. + 200 mg.

3. Add, and express the sum in c.c.:—
$\frac{1}{2}$ litre + 20 c.c. + 200 decilitres

4. A tin measures 150 mm. high, 12 cm. broad, and 24 cm. deep. How many kilograms of water will it hold?

5. If 1 metre of wire weighs 120 centigrams, what will be the length of 60 milligrams of the same wire?

6. Write out a list of the weights in a box where the least is 0·01 gram and the greatest 100 grams.

7. Draw a portion of a burette stem including the figures 10, 11, and 12, and show the surface of water within it standing at 10·18 c.c.

8. Describe fully an experiment which shows that a liquid expands when heated.

9. Write a description of a mercurial thermometer. What numbers mark the "fixed points" on a Fahrenheit thermometer?

10. Draw the apparatus used to mark the boiling point on a thermometer.

11. Find (*a*) the C. readings corresponding to 122° F. and 14° F.
(*b*) ,, F. ,, ,, ,, 20° C. and −5° C.

12. Describe some way in which the volume of a glass stopper could be found.

(*b*)

13. If one end of a straw is dipped into some lemonade and you suck at the other end the lemonade ascends the straw. How do you account for this?

14. Describe in detail how to set up a mercurial barometer. What will be its usual height in inches?

Why does the height of a barometer change sometimes,

(*a*) in the course of a railway journey?
(*b*) at the same place in the course of a few hours?

15. What differences are there between the two flames which can be obtained with a Bunsen burner? State exactly what kind of flame and what part of the flame you would use

(*a*) to bend a glass tube,
(*b*) to smooth the sharp edge of a glass tube.

16. A beaker is full of powdered ice and a thermometer is plunged in it. State fully

(*a*) what changes you will see,
(*b*) what temperatures the thermometer will show if a flame is left under the beaker until all the ice is melted, and the water formed finally boils.

17. Explain what is meant by the statement that when certain substances are heated "physical changes" may take place. Give one example.

18. Draw a Liebig condenser in section.

(*c*)

19. Find (*a*) the F. temperature corresponding to 18° C.
(*b*) ,, C. ,, ,, ,, 23° F.

20. How have you proved that there is a vacuum above the mercury in a barometer tube?

Why does the mercury stand so high? Why does it not reach up to the top of the tube?

21. Describe what occurs on heating

(*a*) tin in an iron spoon,
(*b*) mercury in a glass test-tube,
(*c*) olive oil in a flask.

22. Explain fully why the tumblers of water on the dinner table sometimes become bedimmed with dew.

23. Define or explain the terms—

(*a*) the "melting-point" of a solid,
(*b*) the "solvent action" of a liquid,
(*c*) the "latent heat of vaporisation" of a liquid.

24. Draw carefully a wash-bottle.

25. What do you understand by the terms "hard" and "soft" as applied to water? What would be the effect on the "hardness" of a "hard" water of—

(*a*) shaking it with washing soda,
(*b*) shaking it with gypsum,
(*c*) boiling it,
(*d*) distilling it?

26. 10 c.c. of a solution of salt were found to weigh 11·86 grams, and to leave 3·05 grams residue after evaporation. Calculate the amount of salt dissolved in 100 grams of water.

(*d*)

27. Define the "relative density" of a substance?
You are provided with an irregular lump of granite. Describe in detail two methods by which you could find its relative density. Which will give the more accurate result?

28. A glass stopper weighs 60 grams in air, 35 grams in water, and 40 grams in spirits. Find the relative density of the spirits.

29. A clear liquid is said to be pure water. Describe all the experiments you would make in order to prove whether it is so or not.

30. You are provided with some surface soil, which contains a small quantity of nitre. Describe in detail how you would obtain from it pure crystals of nitre.

31. What is "water of crystallisation"? Instance two kinds of crystals which contain it, and two kinds which do not.

32. Calculate the solubility of potassium chlorate at the ordinary temperature if 16·14 grams of solution on evaporation are found to contain 1·01 gram of the substance.

(*e*)

33. Plot the following barometric heights in pencil on squared paper:—

June 1	30·16 ins.	June 9	30·35 ins.
,, 2	29·94 ,,	,, 10	30·25 ,,
,, 3	30·24 ,,	,, 11	30·25 ,,
,, 4	30·14 ,,	,, 12	30·11 ,,
,, 5	30·18 ,,	,, 13	30·13 ,,
,, 6	30·23 ,,	,, 14	29·99 ,,
,, 7	30·30 ,,	,, 15	30·05 ,,
,, 8	30·35 ,,		

34. Explain clearly how it is that (*a*) *sunshine*, and (*b*) *wind* increase the rate at which a wet road dries.

35. When a long frost is followed by a few warm days the walls inside a house are often found dripping with moisture. Account for this.

36. You are provided with some muddy water. Describe in detail how you would prepare some pure water from it.

37. Define a *calorie*, and the *specific heat* of a metal.

Calculate the specific heat of a metal if :—

Weight of metal	25 gm.
Volume of water	50 c.c.
Temperature of hot metal	100° C.
,, cold water	11° C.
,, water after experiment	15° C.

38. Draw carefully the apparatus which was used to measure the latent heat of vaporisation of water.

39. If the L.H. of liquefaction of ice is 80, and the L.H. of vaporisation of water is 540, calculate the amount of heat required to change 1 gram of ice at 0° C. into steam at 100° C.

(*f*)

40. Four experiments were made to find the percentage of water in a crystal :—

(*a*) 2·64 grams lost 0·65 when heated
(*b*) 2·97 ,, ,, 0·75 ,, ,,
(*c*) 2·33 ,, ,, 0·58 ,, ,,
(*d*) 2·06 ,, ,, 0·52 ,, ,,

Calculate in each case the percentage loss in weight, and also the *average* result.

41. If you were given a muddy liquid containing a soluble and an insoluble constituent, describe in detail the steps you would take :—

(*a*) to separate the *whole* of the soluble portion,
(*b*) to obtain a specimen of the soluble portion in *crystals*, if possible.

42. Describe how you would distinguish between :—

saltpetre and sal-ammonia,
caustic soda and washing soda,
sulphuric acid and nitric acid.

43. Draw the apparatus required for the preparation of hydrochloric acid.

Name the chemicals used.

44. What do you call the substance which can be made from hydrochloric acid and caustic soda ?

If the acid was in strong solution, and the soda in solid sticks, explain in detail how you would prepare a sample of the pure compound free from any admixture of the original substances.

45. What are some of the properties of sulphuric acid ? What is the effect of warming it with some nitre crystals ?

PRACTICAL EXAMINATIONS

The following practical tests have all been terminal examinations which have been actually set in schools. The average age of the Forms examined was not less than $13\frac{1}{2}$ years nor more than 15 years.

(*a*) 1 Hour

1. Make a list of, and add up the weights on the pan of the balance.

2. Read the level of water in the burette.

3. Read the temperature of the running water.

4. Measure the distance between the 1st and 10th lines on your paper
   (*a*) in inches and decimal,
   (*b*) in centimetres and decimal.

5. Add water to the measuring jar so that it contains exactly 86 c.c.

6. Weigh the glass stopper.

7. Weigh the beaker: then measure into it 5 times by means of the pipette 100 grains of water. What does the water weigh? Calculate what 1 grain of water weighs.

(*b*) 1 Hour

8. Read (*a*) the level of the liquid in the burette,
   (*b*) the volume of water in the measuring jar.

9. Select a cork to fit the test-tube. Fit it with a glass tube bent twice at right angles.

10. You are provided with a substance A.[1]
    (1) Describe its appearance.
    (2) Describe the effect of heating a small quantity in a dry test-tube.
    (3) Find whether it is soluble in water. State how you proceed.
    (4) Find the effect on the temperature of 100 c.c. of water of adding 30 grams of A to it, and stirring.

[1] AmCl.

(*c*) 1½ Hours

11. Find the relative density of the glass stopper.

12. Describe the appearance of B.[1]
Heat some of B in a dry test-tube. Describe in detail what you observe.

13. Find the percentage loss in weight when about 2 grams of B are heated in a crucible upon a sand-bath.

14. Find the action of water, both hot and cold, upon B.
Prepare from B as large crystals as the time permits—both clean and dry. Describe your method.

(*d*) 1½ Hours

15. Heat C[2] in a dry test-tube. Describe what occurs.

16. Identify the two solids D and E, and the two liquids G and H.

17. You will be provided with a stick of solid caustic potash. Prepare a good specimen of the salt which can be obtained from it and sulphuric acid.

(*e*) 1½ Hours

18. You are provided with a solution of a certain substance[3] :—

(*a*) Find the relative density of the solution.
(*b*) Prepare by crystallisation some of the dissolved substance. Leave a specimen for inspection.
(*c*) Describe the appearance of the crystals obtained.
(*d*) Heat some of the crystals in a dry test-tube, and describe in detail what you observe.
(*e*) Find the percentage loss in weight when about 2 grams of the crystals are heated in a crucible upon a sand-bath.

(*f*) 1½ Hours.

19. Find the percentage loss in weight when about 2 grams of K[4] are heated in a crucible.

20. What is the action of water upon K? State what experiments you try.

21. Separate from K a sample of each of the two substances which it contains. Describe your method.

---

[1] Powdered $ZnSO_4 . 7H_2O$. [2] HgI_2. [3] $ZnSO_4 . 7H_2O$.
[4] 50 per cent. SiO_2 + 50 per cent. $CuSO_4 . 5H_2O$.

ANSWERS

Examples I.

Metric Prefixes: Length. (Page 11.)

1. 50 cm.
2. 100 mm.
3. 10,000 m.
4. 1,000,000 mm.
5. 0·1 dm.
6. 0·000001 Km.
7. 4·56 dm.
8. 234,000 cm.
9. 0·9 cm.
10. 12,345·6 cm.
11. 40.
12. 5000.
13. 4 cm.
14. 2 Km.
15. 120 cm.
16. 4.
17. 40,000.
18. 4.

Examples II.

Metric Square Measure: Area. (Page 14.)

1. 100 sq. dm.
2. 10,000 sq. cm.
3. 60 sq. cm.
4. 24 sq. cm.
5. 625 sq. mm.
6. 100 sq. cm.
7. 64.
8. 25,000.
9. $196\frac{3}{7}$ cwt.
10. $1\frac{19}{81}$ sq. yds.

Examples III.

Metric Cubic and Capacity Measure. (Page 18.)

1. 1000 c.c.
2. 27,000 c.c.
3. 15,625 cu. mm.
4. 1000 cu. mm.
5. 1,000,000 c.c.
6. 3000 c.c.
7. 200 c.c.
8. 1·234 litre.
9. 56,780 c.c.
10. 1200 c.c.
11. 600 cu. mm.
12. 4320 c.c.
13. 2 litres.
14. 400.
15. 324,000 litres.
16. $8\frac{1}{3}$ days.

Examples IV.

Metric System: Weight. (Page 19.)

1. 10 dgm.
2. 10 mgm.
3. 7·891 Kgm.
4. 0·2 gm.
5. 0·05 gm.
6. 6540 gm.
7. 0·1 gm.
8. 34 mgm.
9. 250 mgm.
10. 3·25 gm.
11. 25·202 gm.
12. 1023·4 gm.
13. 25 c.c.
14. 27 gm.
15. 1150 gm.
16. 8000 gm.
17. 1350 gm.
18. 55 c.c.
19. 5 gm.
20. 1 gm.

Examples V.

Metric System: Miscellaneous. (Page 21.)

(*a*)

1. 20.
2. 50.
3. 20 mgm.
4. 1370 gm.
5. 2777·7̇ cm. per sec.
6. 0·01 gm.

(*b*)

7. 1450·632 m.
8. 100 sq. cm.
9. 4·32 cu. dm.
10. 0·4 gm.
11. 0·5 gm.
12. 30·48 cm.

(*c*)

13. 1920 c.c.
14. 8075.
15. 6 gm.
16. £125.
17. 5 m.
18. 23·7 Km.

(*d*)

19. 2400·71 gm.
20. 100 gm.
21. 20 times.
22. 50 gm.
23. 1370 gm.
24. 37·0̇37̇ ounces.

Examples VI.

Burette Reading. (Page 35.)

1. A. 27.
   B. 43.
   C. 6·7.
   D. 830.
   E. 114.
   F. 136.
   G. 235.
   H. 265.

2. I. 1·68.
   K. 13·32.
   L. 10·68.
   M. 23·2.
   O. 46·8.
   P. 632.
   Q. 868.
   R. 680.

Examples VII.

Interpolation: Conversion of Thermometric Scales. (Page 46.)

1. 63.
2. 1·31.
3. 5·58$\frac{4}{7}$ p.m.
4. 1·361.
5. 77° F.
6. 41° F.
7. 176° F.
8. 10° C.
9. 35° C.
10. 50° C.
11. 68° F.
12. 23° F.
13. 40°·1 F.
14. 4°·4 C.
15. − 10° C.
16. − 20° C.
17. 12° R.
18. 68° F.

Examples VIII.

Relative Densities of Liquids. (Page 53.)

(*a*)

1. 147 gm.
2. 800 gm.
3. 0·862.
4. 1·023.
5. Water.
6. ·074 c.c.
7. 1111·1 c.c.
8. 1·25 c.c.
9. 50 c.c.
10. 25 c.c.

(*b*)

11. 13·5 ; ·07407.
12. 38·09 c.c.
13. 70·85 gm.
14. 36·2 gm.
15. ·940.
16. 1·026.
17. 60 c.c.
18. 150 c.c.
19. 250 c.c.
20. ·923.
21. 2800 gm.
22. ·98.
23. 20 sq. mm.

Examples IX.

Relative Densities of Solids. (Page 57.)

1. (1) 8·95.
   (2) 7·2.
   (3) 0·25.
   (4) 11·4.
2. 2·7.
3. 0·6.
4. 2·5 ; 2·14 ; 10.
5. 8.
6. 22·8 gm.
7. 179$\frac{5}{164}$.
8. £0 0*s.* 11$\frac{1}{4}$*d.*
9. 50 oz.
10. 0·001 c.c.

EXAMPLES X.

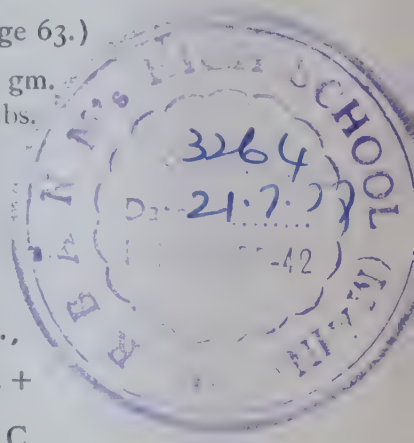
Relative Densities of Solids. (Page 59.)

1. 2·5.
2. 7·5.
3. 11·4.
4. 145 gm.

EXAMPLES XI.

Relative Densities of Solids. (Page 63.)

1. 17·5.
2. 2·72.
3. 2·00.
4. 2·62.
5. Copper: 8·91.
6. 37·6 gm.
7. 7 7 lbs.
8. 2·4.
9. 0·8.
10. 8·6.



EXAMPLES XII.

(*a*) *Specific Heat.* (Page 117.)

1. i. 33°·5 C.
   ii. 32°·41 C.
   iii. 25° C.
   iv. 56°·8 C.
2. i. 475 calories.
   ii. 403·75 ,,
   iii. 60 ,,
   iv. 8800 ,,
3. 0·057 +.
4. 0·034 +
5. 0·5.
6. 3°·75 C.
7. 0°·94 + C.
8. 51°·2 + C.
9. 2°·9 + C.
10. 154·3 + gm.
11. 333·3 gm.

(*b*) *Latent Heat.* (Page 118.)

12. 80,000 calories.
13. i. 78·4.
    ii. 78·5.
    iii. 80·6.
14. 65°·4 C.
15. 186·2 gm.
16. i. 520.
    ii. 515.
17. 717 calories.

EXAMPLES XIII.

Solubility Curves. (Page 180.)

1. 16, 54, 74 gm.
2. 16°, 52°, 95° C.
3. (i.) 19 gm.
   (ii.) 19 gm.
4. 77° C.
5. (i.) Nitre.
   (ii.) Salt.
6. 1 gram per degree.
7. (i.) 14°·5 C.
   (ii.) 61° C.
8. (i.) 200 gm.
   (ii.) 50 gm.
9. 33 gm.

Examination Papers. (Page 194.)

2. 1000·75 gm.
3. 20,520 c.c.
4. 4·32 Kgm.
5. 5 cm.
11. (*a.*) 50° C. and − 10° C.
    (*b.*) 68° F. and 23° F.
19. (*a.*) 64°·8 F.
    (*b.*) − 5° C.
28. 2·5.
32. 6·6.
37. 0·094.
39. 720 calories.
40. (*a.*) 24·62 per cent.
    (*b.*) 25·25 ,,
    (*c.*) 24·89 ,,
    (*d.*) 25·24 ,,
    Average = 25·00 per cent.

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CHAPTER I

THE DISCOVERY OF "FIXED AIR"[1]

Is Atmospheric Air the only Kind of Air?—There are, as we well know, many kinds of solids, which differ from one another in appearance, in relative density, in melting point, in specific heat, in latent heat of fusion, and in many other properties. Nor are we less aware of the existence of different kinds of liquids.

Are there also different kinds of airs? We know that we live in a vast ocean of air, invisible, but around and about us on every hand. Is this atmospheric air the only kind of air? Or are we unaware of other airs, only because they are invisible, formless, and vague? In view of the fact that there are so many different solids and liquids, it would be very remarkable if there were but one kind of air. We must at least admit the possibility of the existence of different airs, and it is very desirable to follow up any line of inquiry which promises to lead to the discovery of any air different from that we breathe.

The Discovery of "Fixed Air."—As a matter of history, an air differing from atmospheric air was discovered by a Dutch chemist *Van Helmont*, who died in 1644. He gave it the name of **gas**. This word, now so familiar to us, was coined by him for the new air, and is undoubtedly connected with our word *ghost*, and the German word *geist*, a spirit.

His discovery and experiments were, however, forgotten, and it was left for *Black*, Professor of Chemistry at the University

[1] Also called *carbonic acid*, *carbon dioxide*.

of Edinburgh, to establish in 1755 the existence of another air, differing essentially from atmospheric air. Let us now perform an experiment, which will give us the necessary clue to the discovery of this air.

EXPT. 1. **The Action of Acids upon Chalk.**—Place a small piece of chalk in a test-tube. Pour upon it a few drops of hydrochloric acid. A violent effervescence, probably due to the escape of an air or gas, results. Repeat this treatment with marble. There is a similar effervescence.

EXPT. 2. **To find whether a Gas is evolved by the Action of Acids on Chalk.**—If a gas is really given off, bubbles will probably be seen if it is led through water. Test this point as follows :—

Take a 4 oz. flask, and introduce a few grams of chalk (not powdered). Bend a glass tube, as represented in Fig. 1. Attach it as a delivery tube to the flask by means of a cork. Let the end dip below the surface of some water in a tumbler. Now remove the cork. Pour in about 50 c.c. of weak hydrochloric acid, and at once replace the cork. There is a violent effervescence, and at the same time bubbles are forced up through the water.

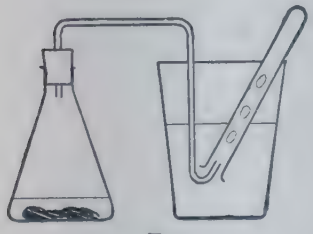

FIG. 1.

Can we collect the gas? Fill a test-tube with water. Close the end with the thumb and invert the test-tube. Place the mouth of the test-tube below the surface of the water in the tumbler, and then remove the thumb. The test-tube remains filled with water. Move the test-tube so that the end is brought above the end of the delivery tube. Bubbles rise up the test-tube, and displace the water. Repeat with *marble*.

There can no longer be any doubt that a gas is evolved by the action of hydrochloric acid on chalk or marble.

EXPT. 3. **Collection of the Gas obtainable by means of Marble and Hydrochloric Acid.**—If we are to examine this gas, we must prepare several samples of it. The apparatus used in the last experiment has the disadvantage that more acid cannot be added without removing the cork, and so admitting also common air. This can be avoided by passing a funnel through the cork.

(1) Place a few lumps of marble (not powdered) in a thick glass bottle. Fit it with a cork, and pass through this a thistle-funnel, *a*, and a delivery tube, *b*, bent as in Fig. 4. Let the end dip below the surface of some water in a trough or basin. (A two-holed rubber stopper may be used with advantage.)

(2) Fill a glass jar completely with water. Cover the open end with a ground glass plate, and see that no air is included. Holding the glass plate on with the right hand, grasp the jar with the left hand, as in Fig. 2.

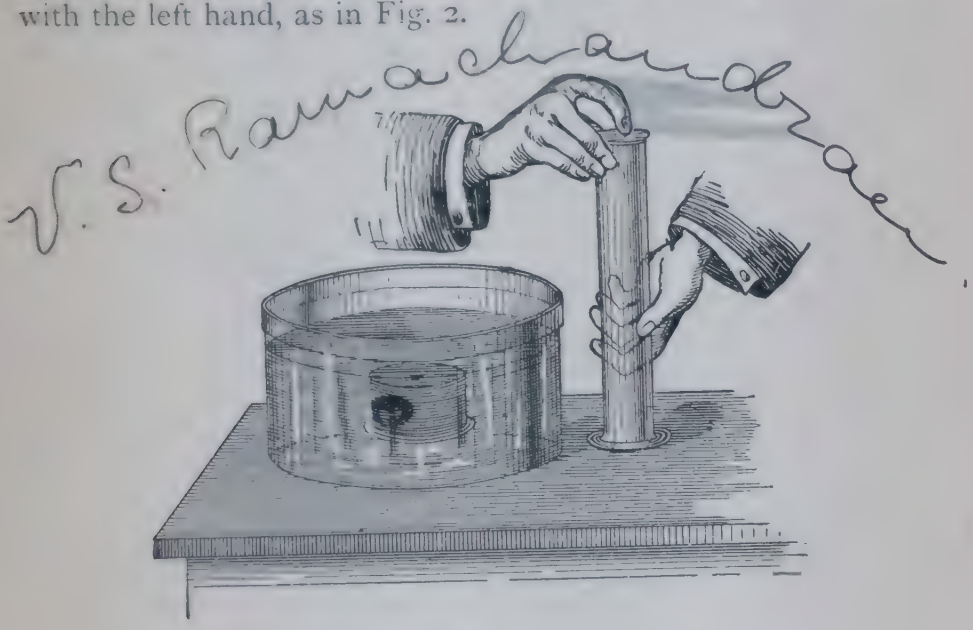

FIG. 2.

(3) Then invert the jar, and place its mouth (still closed by the glass plate) below the surface of the water in the trough, as in Fig. 3.

Now withdraw the plate and allow the jar to stand on the bottom of the trough. Fill three other jars in the same way.

(4) Pour water down the thistle-funnel until the marble is covered. Measure out 50 c.c. of strong hydrochloric acid. The commercial quality will do. Pour some of the acid down the funnel. Bubbles come up through the water.

(5) Now place a jar on a little earthenware "beehive-shelf" above the end of the delivery tube, as in Fig. 4. As soon as the water in the jar has been almost entirely displaced, slide it quickly to one side, slip a glass plate over the mouth (still below the

surface of the water), remove the jar, holding the plate in position, and place it on the bench. At once place another jar over the delivery tube. Collect in this way five or six samples of the gas, and leave these standing mouth downwards on the wet glass plates. Put away the first jar which was filled. (Why?)

(6) Pour water down the thistle-funnel so as to stop the action. Pour the liquid away, leaving the marble in the bottle, and wash it once or twice with water.

This method of collecting the gas is spoken of as **collecting over water.** It is a very convenient method, devised by

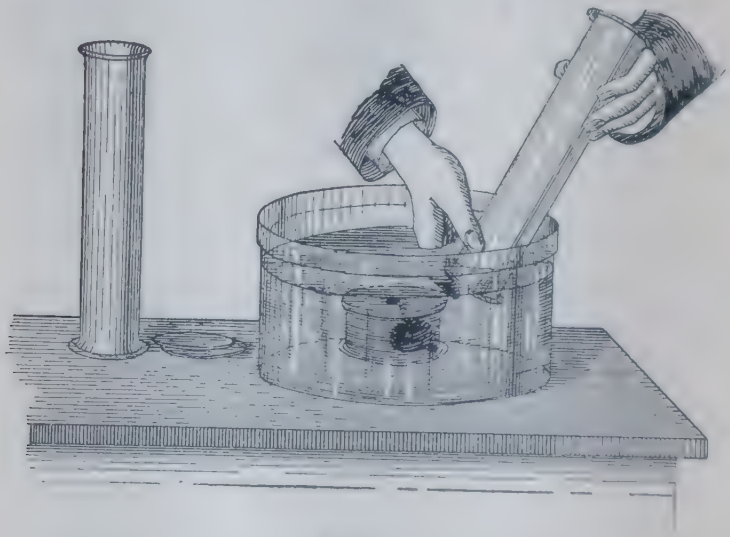

FIG. 3.

Priestley, a chemist who made great use of it in the latter half of the eighteenth century. The trough or basin used in this method is often called the **pneumatic trough** (Greek *pneuma*, air or breath).

Black called the air or gas we have obtained from chalk and hydrochloric acid, "**Fixed Air,**" because it appeared to be fixed in chalk in the solid state. Black said "I have given the name of *fixed air*, and perhaps very improperly: but I thought it better to use a word already familiar, than to invent a new name, before we be more fully acquainted with the nature and properties of this substance which will probably be the subject of my further inquiry." The name is so appropriate, as in-

dicating the preparation of the air by means of a *solid* substance and hydrochloric acid, that we shall use it until we find reason for a better name.

But is Fixed Air after all only common air, or is it a different kind of air or gas? Has it the properties of common air?

THE PROPERTIES OF FIXED AIR

EXPT. 4. **To find whether a Splint or Charcoal will burn in Fixed Air.**—(1) Turn one of the jars previously filled with Fixed Air mouth upwards. Remove the plate, and plunge a lighted splint into it. The flame is immediately extinguished,

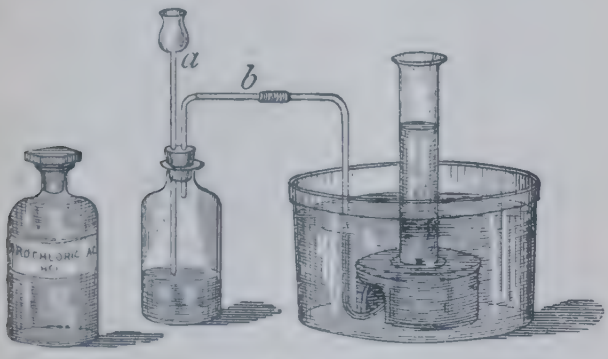

FIG. 4.

and the glowing part of the splint becomes black instantly. Is the splint injured? Remove it and apply a flame. It is rekindled. Plunge it into the jar again. It is once more extinguished.

(2) Bind some wire round a piece of charcoal. Heat it red hot in a flame, and at once plunge it into a jar of the gas. The glowing charcoal becomes black at once, as though plunged into water.

Here, then, we find a difference between this gas and ordinary atmospheric air. And, further, were we to make the experiment, we should find that **no animal could live in this gas.** It would be suffocated.

EXPT. 5. **To find whether Fixed Air is lighter or heavier than Common Air.**—If it is lighter, it will probably float above ordinary air; if heavier, it will sink. Remove the lid

from a jar, A, of the gas (Fig. 5). After one minute insert a burning splint. It goes out. Now hold the jar upside down for a minute in the position B. Introduce the burning splint again. It still burns. So the gas which put out the splint has fallen out of this jar, and is *heavier* than air.

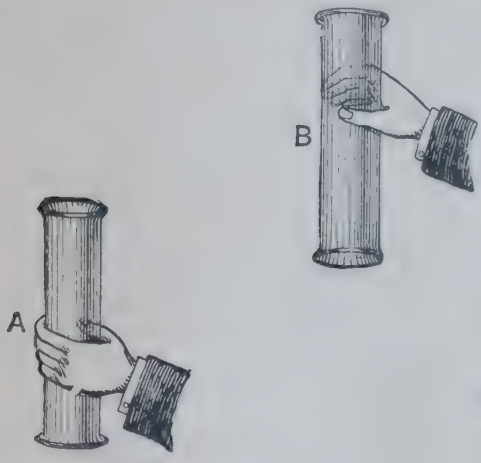

FIG. 5.

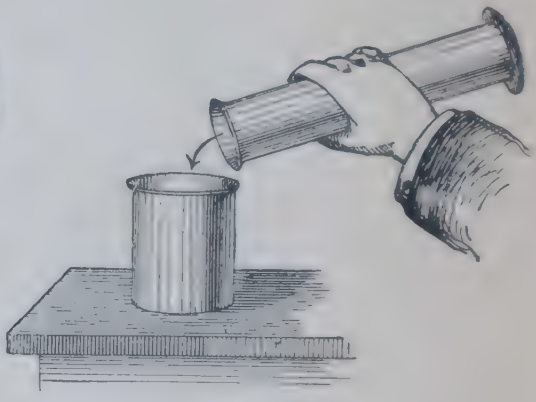

FIG. 6

PROBLEMS

1. Try whether you can pour the gas downwards from a full jar into a small beaker full of ordinary air, as in Fig. 6. Can you show the presence of Fixed Air in the beaker?

2. Try the effect of pouring the gas downwards on to a lighted candle or splint.

3. Try whether you can *syphon* the gas out of a jar down on to the flame of a candle.

The following is another effective way of showing that Fixed Air is heavier than common air.

EXPT. 6.[1] Suspend two thin narrow beakers at the two ends of the arms of a balance. Counterpoise them exactly by adding a little sand to the lighter arm. Take a jar filled with Fixed Air, and holding it in the position shown in Fig. 7, pour its contents into the beaker A. The beaker A descends. Introduce a splint into A and into C, and observe that it is extinguished in A, but burns in C.

[1] Lecture Table Experiment.

Expt. 7. **Water can be charged with Fixed Air.**—Fill an 8 oz. flask with water and invert it in water. Pass Fixed Air into it from a jar of the gas until about half full. Then placing the hand on the upper part of the flask, shake it briskly. In a few minutes the water absorbs the Fixed Air, and taking its place nearly fills the flask as at first. Water dissolves about its own volume of Fixed Air.

To find whether the properties of water are altered by Fixed Air, run a stream of the gas through water in a test-tube. Cut off

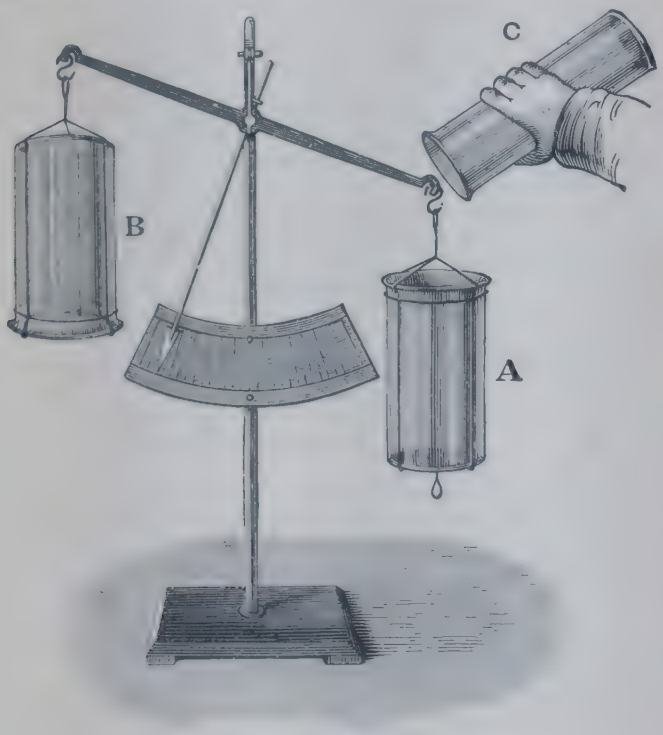

Fig. 7.

the curved end of the delivery tube used in Expt. 3, so that the gas can be led to the bottom of a test-tube half filled with water. After a few minutes taste the water. It has a pleasant faintly acid taste, like soda-water. Try its action on litmus paper; it reddens the paper. Warm the water; it begins to effervesce.

Problems

1. Try the effect of Fixed Air on different liquids. Pass a slow stream of Fixed Air for about half a minute into a few c.c. of several different

liquids placed in turn in a test-tube. Take, for instance, hydrochloric acid, caustic soda, lime-water, very dilute blue litmus solution. Find in each case whether there is any *visible* effect.

2. Try the effect of atmospheric air on the same liquids, producing the current of air by means of an aspirator, or by *sucking* air through the liquid, placed in a small wash-bottle.

3. Try the effect of forcing your breath through a glass tube into the same liquids.

4. Place some clear lime-water in a watch-glass and leave it exposed in the open air. Do you observe any change?

Tabulate your results.

Fixed Air, Expired Air, Atmospheric Air.—Your experiments should show that hydrochloric acid and a solution of caustic soda are not visibly altered by any one of the airs. But dilute blue litmus solution is reddened by Fixed Air, and by expired air, while it is not perceptibly changed by atmospheric air. On the other hand, lime-water is at once made milky by Fixed Air, and very soon also by expired air, and even atmospheric air after some minutes produces the same effect.

Then are Fixed Air, expired air, and atmospheric air one and the same air? Clearly they are not, for we have found Fixed Air and atmospheric air to differ in several respects: Fixed Air extinguishes a flame; it is heavier than atmospheric air; it is slightly acid; it turns lime-water milky much more quickly. The explanation that Fixed Air, expired air, and atmospheric air all produce eventually the same effect upon lime-water seems rather to be that expired air and atmospheric air *contain* a small proportion of Fixed Air. This is, in fact, the case.

10,000 parts of atmospheric air generally only contain 3 parts of Fixed Air. Expired air contains 100 times as much Fixed Air as the atmospheric air breathed into the lungs. This Fixed Air present in expired air is derived from the blood. Fixed Air passes from the blood through the walls of the blood-vessels in the lungs out into the air breathed into the lungs, which is then breathed out again with its increased proportion of Fixed Air. A full-grown man will expire about 14 cubic feet of Fixed Air every day. Fixed Air, as has been already stated, is fatal to animal life, and even when atmospheric air contains no more than 10 parts of Fixed

Air per 10,000, it should be considered far from pure, and it is likely to produce oppression, exhaustion, and headache. Enough, therefore, has already been said to prove the need of proper ventilation in all living-rooms, or in large halls where many people are gathered together.

A Test for Fixed Air.—The experiments which have been made show that Fixed Air is most readily detected by means of lime-water, which is rendered milky by it. Even 3 parts of Fixed Air in 10,000 parts of atmospheric air have been detected.

EXPT. 8. **To show that Chalk and Hydrochloric Acid also give off the same Fixed Air.**—We have found that a gas is given off when hydrochloric acid is added to both marble and chalk. But is it the same gas? Fit a test-tube with a cork and a small delivery tube bent twice at right angles, which will reach to the bottom of a second test-tube, in which some clear lime-water is placed. Place in the tube one or two small pieces of chalk, and cover them with dilute hydrochloric acid. The gas which is given off bubbles through the lime-water, and quickly turns it milky. So chalk and hydrochloric acid liberate the *same* gas, Fixed Air, as marble and the acid.

The Source of the Fixed Air in Experiment 3.—While we have proved that chalk and marble with hydrochloric acid give off a gas which is different from atmospheric air, yet we have no evidence whatever to lead us to conclude whether the Fixed Air comes entirely from the solid, or entirely from the acid, or partly from one and partly from the other, or from both. We cannot at present say which is the case.

But are chalk and marble the only solids which, with hydrochloric acid, will evolve Fixed Air? And is hydrochloric acid the only acid which is capable of doing this?

PROBLEMS

1. Find whether whitening, limestone, calc-spar, sand, plaster of Paris, old mortar, oyster shells, egg shells, soda crystals, baking powder, do or do not give off a gas in contact with hydrochloric acid. If a gas is given off, find in each case whether it is Fixed Air.

2. Find whether chalk in contact with sulphuric acid, nitric acid, or acetic acid gives off Fixed Air.

Laboratory Method of preparing Fixed Air.—Knowing now that Fixed Air is heavier than common air, we can devise a readier way of collecting it than that described in Experiment 3, which is also open to objection because some of the gas disappears into the water.

EXPT. 9. **Collection of Fixed Air by Downward Displacement.**—Attach to a bottle in which Fixed Air is to be produced a glass tube bent at right angles, so that it may reach to the bottom of a jar (Fig. 8). Place about 20 grams of marble in small pieces (not powdered) in the bottle. Add water first, and then an equal volume of hydrochloric acid. The Fixed Air which is evolved will collect at the bottom of the jar and gradually displace all the air. The jar is known to be full of the Fixed Air when a splint held just below its mouth is extinguished.

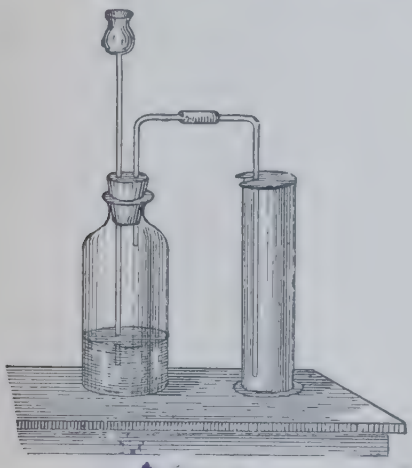

FIG. 8.

This apparatus is to be used in future when Fixed Air has to be prepared.

Summary.—The main result, then, at which we have arrived is that there are several substances which in contact with acids give off a gas which is essentially different from common air, for a splint cannot burn in it; it is heavier than common air; it has an acid character; it renders lime-water milky, and water will dissolve an equal volume of it. This air we have named "Fixed Air."

Black did not leave the matter here, but measured the amount of Fixed Air which could be obtained by means of a given weight of chalk, and subjected chalk itself to a thorough examination.

We shall, for the present, however, leave the study of chalk, in order to follow up other lines of inquiry, which were successfully pursued by a few chemists in the seventeenth and eighteenth centuries. Subsequently, when our hands have greater skill, and our faculties are better trained, we shall

return to Black's famous research, which is a model of what an investigation should be.

Problems

1. Examine the air dissolved in soda-water. Open a soda-water bottle, and pour some of the contents at once into two gas jars. Close the jars with glass plates, and shake. Then test the air in one jar with a lighted splint. To the other jar add some lime-water, and shake.

2. Collect some expired air in two gas jars standing over water, making use of a bent glass tube. Try whether a splint will burn in the air. Test it also with lime-water. What do you observe?

3. Invert a test-tube full of Fixed Air in a dilute solution of caustic soda, and shake gently. What do you observe?

CHAPTER II

THE STUDY OF FIRE AND AIR

THE PHLOGISTIC THEORY OF FIRE.—We have all watched a bright fire, and delighted in the long tongues of flame, flickering and darting upwards, and we know that after it has died out dead ashes alone remain. What has occurred? Is the flame simply something escaping from the coal? Men used to think so. Or, again, take the following instance of another kind of flame.

EXPT. 1. **The Burning of Magnesium.**—Take a strip of magnesium ribbon 10 cm. in length, and holding one end in a pair of tongs, kindle the other end. It burns brilliantly, and a white ash or calx is left. Is this dazzling light the evidence of something parting from the magnesium and leaving the calx?

This idea, the first perhaps that is likely to occur to any one, agrees with one which was put forward in the seventeenth century by Becker, a distinguished chemist, and by Stahl, a physician to the King of Prussia. They taught that all combustible bodies were compounds containing at least two constituents, and that when combustible bodies were burnt, one of these constituents escaped while the other remained.

In the burning of the magnesium, for instance, they would have said that a combustible constituent escaped which had been contained in the metal. They gave the name "phlogiston" to this combustible constituent, and their Theory of Combustion was known as the **Phlogistic Theory**.[1]

An advantage of this theory was that all cases of combustion

[1] They said Phlogiston was *materia aut principium ignis, non ipse ignis*—the substance or principle of fire, not fire itself.

were looked at from the same standpoint. For it was considered that all combustible substances contained the same phlogiston, and in the burning of a candle, an oil-lamp, a fire, sulphur, or a metal such as magnesium, they saw alike the escape of this phlogiston.

THE STUDY OF FIRE IN THE SEVENTEENTH CENTURY.—A result of the neglect of measurement, on the part of the early chemists, was that, although Fire had been an object of wonder and attention since the earliest ages, it was not until the middle of the seventeenth century that any real knowledge of the phenomenon was obtained. This we owe to a little band of chemists in Oxford.[1] It will now be our object, following their guidance, to begin for ourselves the study of Fire and Flame.

EXPT. 2. **To observe the Effect of Heating Wood.**—(1) Place two or three chips of wood in a porcelain crucible, cover it with a lid, and supporting it in a pipe-clay triangle, heat it strongly. Fumes are given off, which may catch fire and burn round the lid. When these cease, remove the flame, allow the crucible to cool, and then examine the contents. The wood has not been all burnt away, but it is left charred and very brittle.

(2) Heat what is left in the crucible, this time without the lid. Watch closely. The charred fragments glow and burn away, leaving only a very light white ash.

Why does the charred wood left in the first experiment burn away in the second? What do these two experiments seem to show?

Robert Hooke, one of the Oxford chemists, gave in 1665 the following description of the method of converting wood into charcoal:—

"The body to be charr'd or coal'd may be put into a Crucible, Pot, or any other Vessel, that will endure to be made red hot in the Fire without breaking, and then cover'd over with Sand, so as no part of it be suffer'd to be open to the Air, then set into a good Fire and there kept till the sand has continu'd hot for a quarter, half, an hour or two, or more, according to the nature and bigness of the body to be coal'd or charr'd, then taking it

[1] See Prof. H. B. Dixon's address before Section B of the British Association, Oxford, 1894.

out of the Fire, and letting it stand till it be quite cold, the body may be taken out of the Sand, well charr'd and cleans'd of its waterish parts." [1]

If, however, whilst the crucible is still hot, the contents are turned out upon an iron plate so that the body is exposed to the air, Hooke observes that "when it comes into the free air, it will take fire, and readily burn away."

This experiment is very important, for it seems to show that **the burning of the charcoal is dependent upon the presence of air**; indeed, this is the conclusion which Hooke drew, for he said :—

"From the experiment of charring of Coals (whereby we see that notwithstanding the great heat and the duration of it, the solid parts of the Wood remain, whilest they are preserv'd from the free access of the air undissipated) we may learn, that which has not, that I know of, been publish'd or hinted, nay, not so much as thought of, by any; and that in short, is this—that the Air in which we live, move and breathe, and which encompasses very many, and cherishes most bodies it encompasses, is the universal dissolvent of all Sulphureous bodies" (*i.e.* of all combustible bodies).

Now if it be true that the air is necessary for the combustion of any body, it follows that a substance, for example **sulphur**, will not burn in absence of air. This *Robert Boyle* proved in 1660 to be the case by an experiment which it is not easy to repeat. He lowered some sulphur on to a hot iron plate in a vessel from which the air had been pumped, and he found that it would not burn, but on allowing a little air to enter, "divers little flashes could be seen," which were extinguished on exhausting the air again.

THE HEATING OF METALS IN THE AIR

In the course of the seventeenth century a number of chemists directed their attention to the effect of heat upon metals, and very important results were obtained from their observations.

EXPT. 3. **To observe the Effect of Heating Lead in**

[1] Hooke, *Micrographia*. *See* Alembic Club Reprint, No. 5, p. 43.

the Air.—Put some lead in an iron spoon.[1] Heat it over the Bunsen flame. After melting, an earthy-looking scum or "*calx*" forms upon the surface of the molten metal. Scrape the calx to one side with an iron wire. A mirror-like metallic surface is exposed, which quickly becomes coated over again. (See Vol. I., p. 84.)

In what way may this calx have arisen? Two causes readily suggest themselves. Firstly, it is possible that something has been driven out of the metal by the heat, leaving the calx as a residue. This is the view which was taken by many chemists 200 years ago.[2] Secondly, it is possible that under the influence of the heat some portion of the air becomes attached to the metal, whereby the calx is formed. Now it is evident that we can decide between these two cases by means of the balance. For if something has been driven out of the metal there must be a *loss* of weight, while if any has been added to it, there must be a *gain* in weight. Let us, therefore, put the matter to the test, and weigh the metal before and after it is heated.

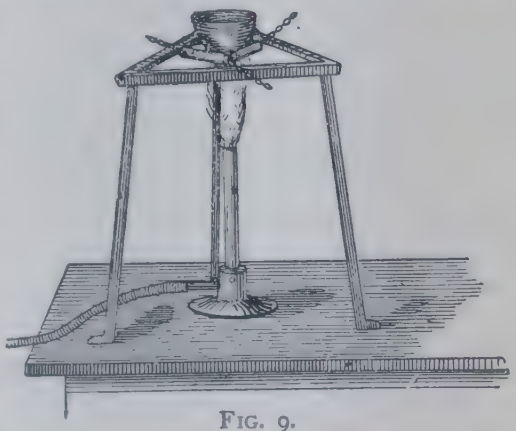

FIG. 9.

EXPT. 4. **To find whether Lead gains Weight when Heated in the Air.**—Weigh a half-ounce crucible. Weigh about 3 grams of clean lead parings into it. Support the crucible on a pipe-clay triangle upon a tripod stand (Fig. 9). Heat it (uncovered) for about twenty minutes over a Bunsen flame. It is well to scrape the calx to one side from time to time with an iron wire, so as to expose a fresh surface, taking care, however, that none is removed from the crucible. Allow it to cool. Weigh the whole again. Is there any increase in

[1] In the absence of sufficient spoons, crucible lids supported on home-made pipe clay triangles can be used

[2] This view was taken by the Phlogistic chemists. They believed the heat drove a fiery material, *phlogiston*, from the metal, and therefore they regarded the calx as a simpler substance than the metal. Their view can be expressed thus:—

calx of metal + phlogiston = metal.

weight? Yes, perhaps 0·03 gram. Repeat. Is there any further increase in weight? Yes. If the calx formed on the surface of the lead is red, it is called *red-lead*. If yellow, it is called *litharge*.

It is so important to decide whether there is an increase in weight when a metal is heated in the air that experiments with other metals must be made.

EXPT. 5. **To find whether Magnesium gains Weight when heated in the Air.**—Weigh a half-ounce crucible with its lid. Weigh 50 centimetres of narrow magnesium ribbon or wire on the pan bearing the crucible. Roll the ribbon up into a loose ball, and place it in the crucible. Support the crucible with the lid upon a pipe-clay triangle on a tripod stand. Then apply a flame until the bottom of the crucible is red-hot. This should take five minutes. Continue heating strongly for another five minutes, but lift the lid about every half minute *very slightly* so as to let in air but not to allow the escape of fumes. Allow to cool. Weigh again. Is there an increase in weight? Yes. Calculate the percentage gain in weight.

EXPT. 6.[1] **To find whether Iron gains Weight when heated in the Air.**—Suspend a horse-shoe magnet from the hook of a balance pan as shown in Fig. 10. Place a round of filter-paper on the pan below. On the paper place a little heap of finely divided iron.[2] Lift the paper and allow the magnet to pick up as much iron as possible. Weigh together the magnet, iron, and paper. Let the flame of a Bunsen burner play gently against the iron on the magnet for a moment. The iron begins to glow. The glow continues and spreads. The iron is burning. Let it cool. Weigh again. It is heavier.

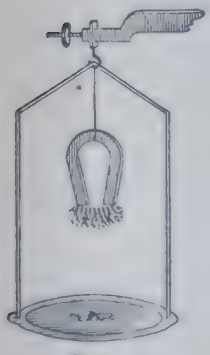

FIG. 10.

In the seventeenth century experiments were frequently made upon the calcination of **antimony** by means of the sun's rays concentrated upon it by a powerful burning glass. A small cone of the metal was weighed and placed upon a slab of marble, and the point of the refracted

[1] Lecture Table Experiment.

[2] *Note for Teacher.*—Iron filings are not satisfactory as finely powdered Commercial iron. If spoiled or rusty, treat the iron in a fresh [illegible] coal gas over a Ramsay burner. Wait until quite cool before removing it.

rays directed on to the point of the cone. The metal immediately fumed, and was converted into a white calx, which was separated with a knife, and the rays directed on to the rest till the whole became white. It was then weighed again. As a writer said in 1630:—

"It is a remarkable thing that although in this calcination the antimony has lost much of its substance, by the vapours and fumes which are copiously exhaled, yet its weight augments instead of diminishing."

From what source comes this increase in weight, when a metal is calcined? In view of all the experiments we have made, it is scarcely possible to conclude that the increase arises from anything else but the *fixing of particles of air in the metal during the heating*. We can scarcely believe that the increase is simply due to heat itself having weight, because Experiment 2 clearly showed that the presence of air was necessary for the burning of charcoal.

THE BURNING OF PHOSPHORUS.—The point at which we have arrived is that air is necessary for the combustion of charcoal, and that when a metal is heated it increases in weight, probably because it combines with particles of air. But can we prove that this is so? Perhaps by heating a substance in a *closed* vessel. If this is then opened under water, water will rush in to take the place of any air which has disappeared.

It is not easy to test this in the case of metals. But in the year 1674 a very remarkable substance called **phosphorus** was discovered, which is particularly suitable for experiments upon this subject.

Phosphorus can be prepared in two very different forms: one is a yellow waxy substance, the other a dark red powder.

It was first prepared from urine by an alchemist named Brand in Hamburg by a secret process. A few years afterwards it was discovered how to prepare it from bone ashes. Even at the present time the details of its preparation are kept secret by the few firms which are concerned with its manufacture.

Its properties are so remarkable that it was exhibited as one of the Wonders of Nature to King Charles II.

EXPT. 7. **To observe the Burning of Phosphorus in the Air.**—Obtain a gas jar. Cut a clean piece of phosphorus, about the size of a pea, and dry it with blotting-paper. Place it in a "deflagrating spoon." Warm one end of a glass rod in a flame, and then touch the phosphorus with it. The phosphorus catches fire. Introduce it into the jar. Dense white fumes are evolved, and presently the flame dies out. Remove the spoon. Some phosphorus remains unburnt, so perhaps it can be re-kindled. Try. Yes, it will burn further in the air outside the jar. Examine the side of the jar. It is covered with a white powder.

How has the white powder arisen? Has a fiery material been driven out of the phosphorus, leaving the powder as a residue? Or has some portion of the air become attached to the phosphorus, so forming the white powder?

We can much more readily try whether any air disappears when phosphorus is burnt than when a metal is burnt or heated in the following way.

EXPT. 8. **Does some of the Air in a Closed Flask disappear when Phosphorus is burnt in it?**—Take a clean and dry 8-oz. flask.[1] Select a tightly-fitting rubber stopper. Cut a clean piece of phosphorus,[2] about the size of a pea, and dry it with blotting-paper. Place it in the flask and close it *firmly* with the stopper. Weigh the whole. Then, holding[3] the flask by the neck, and away from your face, warm it everywhere except where the phosphorus is. Turn the flask and let the phosphorus fall on to its heated sides. It burns in the enclosed space (probably without bursting the flask, but beware!). Replace the flask on the pan of the balance. It may appear lighter than before. (Why?) Allow it to cool. *Its weight has not been changed*, showing that the heat which has passed through the walls of the flask weighs nothing. Now place the neck of the flask below the surface of some water, and

[1] A flask can be dried by waving it over the flame of a Bunsen burner and sucking out the air by a glass tube placed in the flask. Allow it to cool completely before weighing it.

[2] *Caution.*—Phosphorus is *kept* under water, and *cut* under water. It is handled with tongs. Before burning it is dried by pressing between folds of filter-paper. It catches fire of itself, at a temperature below that of boiling water. Let it cool and burn in absence of air. If the hands are burnt by phosphorus, cover the wounds with a mixture of olive oil and slaked lime, called "carron oil."

[3] Hold it for safety within a fume-cupboard.

remove the stopper. Observe that the water rises into the flask ; shake the flask ; the white deposit on the sides of the flask and the fumes are dissolved. Replace the stopper. Remove the flask from the water and dry the outer surface. Measure the water which has entered the flask by pouring it into a graduated cylinder. Measure in the same way the whole volume of the flask. It will be found that the water which enters the flask is just about *one-fifth* of that which it can contain.[1]

It is evident, therefore, that in the combustion of the phosphorus, *one-fifth of the volume of the air disappears.* What has become of it? It cannot have escaped from the flask, for there was no loss of weight. Is it not likely, as was suggested before, that it has united with the phosphorus, forming these white fumes which filled the flask, and which dissolved in the water? (Test the water with litmus paper, and observe that it is acid.)

"But," as Lavoisier, a famous French chemist, said when he was studying the matter, "there is a great difference between conjecture and proof, and it is necessary that it should first be established that a combination of any kind is formed with phosphorus during its combustion."

Now, if the air which has disappeared has really united with the burning phosphorus, the phosphorus must have increased in weight, just as the metals did. In order to make sure whether this is so or not, the burning of phosphorus may be carried out as follows :—

EXPT. 9.[2] **To find whether the Products of Burning Phosphorus are heavier than the Original Phosphorus.**—Take a piece of wide glass tubing about 6 inches long, drawn out at one end[3] (Fig. 11). Place a small piece of phosphorus at D. Pack about 3 inches of it from B to C with fine asbestos. Close the end with a cork through which a glass tube passes. Test by sucking to be sure that air can filter through the asbestos. Weigh the whole. Then fix the tube in a horizontal position, and connect the end B with an aspirator. Ignite the phosphorus, and the fumes will be carried towards B and caught by the asbestos. Finally, when the phosphorus is

[1] The experiment in this form is due to the Swedish chemist Scheele (about 1773 ; *see* Alembic Club Reprint, No. 8, p. 13).
[2] This experiment was suggested by Prof. H. E. Armstrong.
[3] Use a straight "calcium chloride" tube with one bulb.

consumed, weigh the tube again. There is a very decided increase in its weight.

There can therefore be no doubt that, when phosphorus is burnt in a definite volume of air, one-fifth of the air disappears,

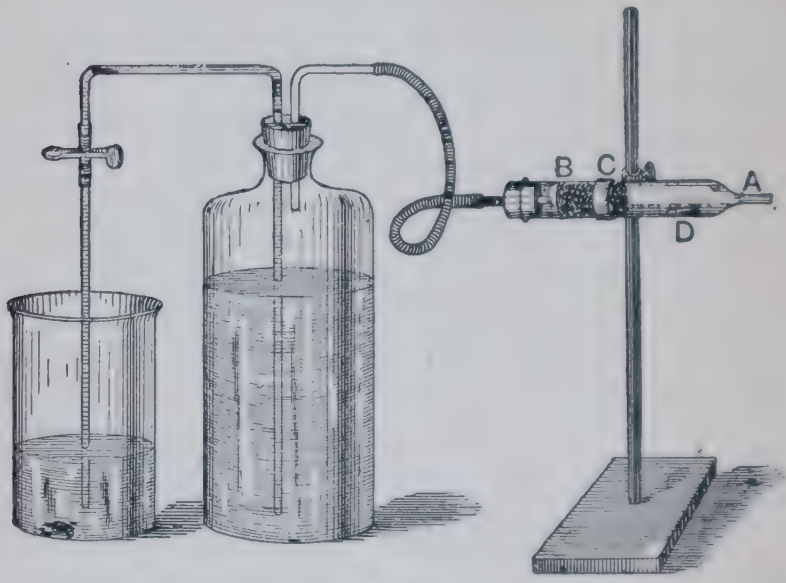

FIG. 11

and it does so because it unites with the phosphorus, forming a white solid substance.

Let us now make two simple experiments with a **CANDLE**.

EXPT. 10. **A Burning Candle needs Fresh Air.**—Attach a candle to a deflagrating spoon. Light it. Put it inside a tall dry glass jar, closing the top of the jar with the plate attached to the spoon (Fig. 12). The flame burns for a time, sinks, and goes out, while the inside of the jar becomes dim. Has anything hurt the candle? Take it out, leaving the jar covered with a glass plate. Try whether the candle can be lighted again. It burns: clearly it is in no way spoilt. Has the air been spoilt? Lower the burning candle into the jar again. It goes out almost immediately. Then it is the air which has been altered by the burning candle.

EXPT. 11. **To find whether Air disappears when a Candle is burnt in a Closed Vessel.**—The following is an

experiment which Mayow showed in Oxford so long ago as 1670. He wrote :—

"Let a lighted candle be so placed in water that the burning wick shall rise about six fingers' breadth above the water ; then let a glass vessel[1] of sufficient height be inverted over the candle. Care must be taken that the surface of the water within the glass shall be equal in height to that without, which may be done by including one leg of a bent syphon within the vessel while the other opens outside. The object of the syphon is that the air enclosed by the vessel and compressed by its immersion into the water may escape through the hollow syphon. When the air ceases to issue, the syphon is immediately withdrawn, so that no air can afterwards get into the glass. In a short time you will see the water gradually *rising* into the vessel while the candle still burns."

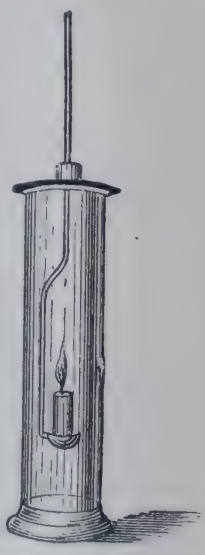

FIG. 12.

Conclusions.—On the evidence brought forward we can now state :—

1. It is the air in which we live that enables combustible substances to burn.

2. When a metal or phosphorus is burnt, it increases in weight.

3. The increase in weight of phosphorus when burnt is due to air becoming fixed in it.

4. When phosphorus is burnt in a closed volume of air, only one-fifth of the air disappears, not the whole.

Observe that it cannot be held that when phosphorus is burnt in a closed volume of air only a part of the air is removed because the rest cannot reach the phosphorus, since the same result may be obtained again and again ; that is to say, only 20 per cent. of the air will be removed each time, *whatever* the size of the flask, and unburnt phosphorus will be left over, if sufficient has been taken.

These facts make it *almost certain that the air is not really a*

[1] A 14-oz. flask with a short neck is suitable. Use a taper in place of a candle, and tie it to a nail forced through a small piece of sheet lead.

single substance, but contains at any rate two gases, one of which alone (occupying one-fifth of the whole) is concerned in combustion ; this portion of the air may conveniently be called "**Fire-Air.**"

THE DISCOVERY OF "SPOILT AIR."—Let us now remove the Fire-Air from a closed volume of common air, by burning some phosphorus in it, and try what kind of air is left.

EXPT. 12. **To remove the Fire-Air from Common Air, and to find whether the Air which is left supports Combustion.** — Take a large glass-bell jar, A (Fig. 13), with a stopper at the neck, and graduate it into six equal divisions.[1] Float a small porcelain dish, D, containing a piece of dry yellow phosphorus in some water in a glass dish, B. Colour the water with litmus solution. Cover it with the bell-jar, so that the water rises to the fifth mark. Touch the phosphorus with a hot glass rod, and quickly replace the stopper. Dense fumes are given off, which soon dissolve in the water, and the water finally rises almost to the fourth mark, if water is poured into the dish so that its level inside and outside the jar is the same. Remove the stopper. Introduce for a moment a lighted candle. It is at once extinguished. Place a small piece of phosphorus or sulphur in the cup of a deflagrating spoon. Ignite it. Introduce it into the jar for a moment. It is extinguished.

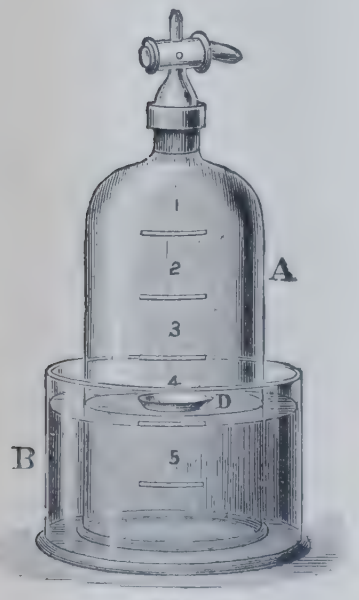

FIG. 13.

It was also found more than 100 years ago that *animals died almost immediately* on being introduced into this portion of common air which was left after the removal of Fire-Air. Let us for the present call this air "**Spoilt Air.**"

Now it is one of the most characteristic of the properties of

[1] The experiment can be made with an ordinary bell jar. In that case some of the gas must be decanted into gas jars for testing.

common air that it is necessary for the life of animals. This has long been known, and it was first clearly proved by Boyle in 1660 by experiments which the student is recommended not to repeat ; he placed cats and mice, sparrows, fishes, and tadpoles in the receiver of his air-pump, and terminated their lives by pumping out the air. It has been stated above that when the Fire-Air has been removed from common air animals cannot live in the air which remains. The conclusion which necessarily follows is that *it is only that portion of the air which we have called Fire-Air which enables animals to breathe* and is concerned in what is called respiration.

Is this Spoilt Air, however, the same as Fixed Air ?

Expt. 13. **To find whether Spoilt Air is the same as Fixed Air.**—To the top of A (Fig. 13) attach a rubber tube into which some cotton-wool is loosely packed. Add more water to the trough. Let the tube dip into some clear lime-water in a test-tube. Open the tap so that some of the air may pass through the lime-water. (The object of the cotton-wool is to filter out all the white fumes.) No turbidity results. Then Spoilt Air must be a different kind of air from Fixed Air, though both of them extinguish a lighted taper.

Exercises

Observe the effect of heating the following metals on porcelain lids. Endeavour to explain the changes effected.

Tin. Zinc. Copper (bright turnings).

Although we have proved that many metals gain in weight when heated in the air, we have really in their case no *direct* evidence that this gain is due to union of the metal with Fire-Air rather than with Spoilt Air. That the union is with Fire-Air is by analogy probable ; but how may we seek for real proof ? Perhaps by passing a current of atmospheric air over some metal heated within a glass tube, collecting the air which passes through, and trying whether it has the properties of Spoilt Air or of Fire-Air. Copper, which is very readily acted upon when heated in the air, with formation of a black scale or calx, is perhaps the most suitable metal to employ.

Expt. 14. **To collect and examine the Air which has passed over Heated Copper.**—Take a hard glass

tube (*ab*) about 12 inches in length. Pack the tube with bright copper turnings, and support it over a Ramsay burner, which will give a long horizontal flame. Fill a large flask (*c*) with water, and support it near the top of an aspirator (*f*). Place in it a glass tube which will reach the bottom of the flask, and connect the other end, by rubber tubing, with the longer tube of the aspirator. Place a screw clamp (*d*) upon this rubber tube. Attach a rubber tube (*e*) to the aspirator, and suck until water is syphoned over from the flask into the aspirator; then

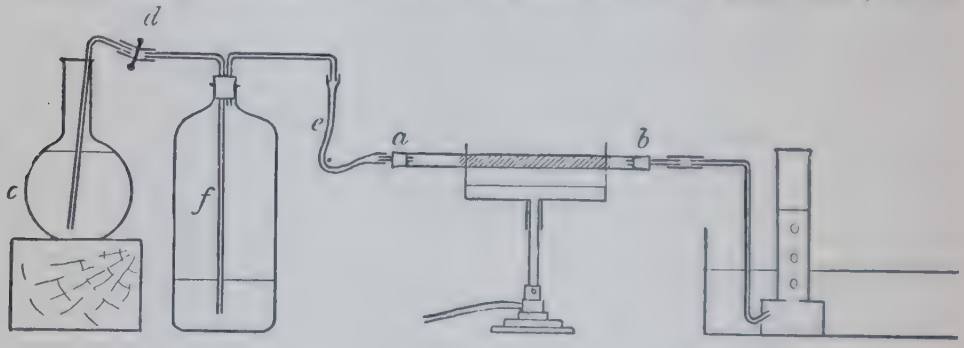

FIG. 14.

tighten the clamp, and connect the tube *e* with one end of the hard glass tube. Attach to its other end (*b*) a short delivery tube, dipping beneath water, below a small gas jar.

Carefully open the clamp until a slow stream of air is forced up into the jar. Then light the burner. Neglect the first few bubbles of air which pass through, and then collect a jar full. Insert a burning candle and a burning splint of wood into the air which has been collected. Both are extinguished.

This experiment, therefore, proves that the metal copper when heated in atmospheric air unites with the *same* part namely, Fire-Air, as does phosphorus.

Common Air is not a Simple Substance.—There can now be no doubt that ordinary air contains at any rate *two airs*. One we have called **Fire-Air**: this it is which enables combustible substances to burn and animals to breathe. The other is entirely different from Fire-Air, for it can neither support combustion nor life. This we have called **Spoilt Air.**

This extremely important conclusion, namely, that common

air contains two entirely different airs or gases, was first arrived at about 1670, from experiments similar to those we have seen, carried out by the Oxford chemist *Mayow*. This discovery, however, attracted little attention, and was soon entirely forgotten. It was left for the Swedish chemist *Scheele* to remake it about the year 1772, and he it was who gave the two airs the names Fire-Air and Spoilt Air.

It must not be forgotten that atmospheric air also contains a varying amount of Water Vapour (see Vol. I., Chapter XI.), and also a very small amount of Fixed Air (see Vol. I., p. 8).

The Nature of Flame and Fire.—Can we now understand the nature of the flame of phosphorus when it is burnt in a closed vessel? Nothing material escapes from the vessel. The original contents were Fire-Air and Spoilt Air. During the experiment light and heat are given out. After the experiment the white dust on the glass is Fire-Air combined with the phosphorus. As Hooke said of an experiment in 1665:—

"It seems reasonable to think that that shining transient body which we call *Flame* is nothing else but a mixture of Air and volatil parts of combustible bodies which are acting upon each other whilst they ascend, and that *this action produces a very great heat*, and that which we call Fire; and that this action is perform'd with so great a violence, and does so rapidly agitate the smallest parts of the combustible matter that it produces in the medium of the Air the action or pulse of Light."

This theory of flame applies equally to the burning of sulphur or of a candle in the air. The heat and light of a burning candle, for instance, we are to regard as the *result* of the violence of the action of the Fire-Air in the atmosphere upon the combustible material of the candle, the heat being so great that the burning materials are heated white hot, and give out a bright light. The flame of a candle, therefore, is the visible evidence of action going on between the Fire-Air and the combustible matter of the candle.

It is this vigorous action between Fire-Air and combustible material, resulting in heat and light, to which the name *combustion* is commonly given, though we shall find afterwards that the word must be used in a somewhat wider sense.

Exercises and Problems

Fire and Air

1. Observe the effect of heating the following substances on porcelain lids :—

Antimony (powdered). Iron (filings). Solder.

2. Observe the effect of holding a thin wire or sheet of the following metals by means of a pair of tongs in the upper part of a Bunsen flame :—

Copper.	Aluminium.
Iron.	Platinum.

3. Find whether there is any change in weight when the following metals are heated in open porcelain crucibles :—

Solder.	Zinc.
Tin.	Copper.

4. Find whether any air disappears when some sulphur, melted and ignited in an iron spoon, is poured into a dry flask, which is then immediately stoppered.

5. Burn a piece of phosphorus in a closed flask as in Experiment 8, p. 18. Weigh the flask. Then remove the stopper, replace it, and weigh again. Is there any change in weight?

6. Find whether the products of burning sulphur are heavier than the original sulphur, drawing the products through a weighed tube containing soda lime.

7. Prepare Spoilt Air as in Experiment 12. Pass it over bright copper turnings within a hard glass tube, first displacing the air and then heating the tube. Is the copper acted upon?

CHAPTER III

THE RUSTING OF IRON

THE RUSTING OF IRON.—In the course of our study of Fire and Air we found that finely divided iron would actually burn in the air, and that in burning it gained in weight. Iron, however, often undergoes some kind of change even at the ordinary temperature; it readily rusts. The bright steel of a bicycle which has been ridden in wet weather is quickly corroded. The rails of a disused railway rapidly acquire a thick coating of red rust.

Take any piece of old iron which has stood in the open air for some time. Scrape off some of the rust and examine it carefully. Have you met with anything like it before in your experiments?

Here then is a problem ready to our hand.[1] What is it that occurs in such cases to iron or steel? We cannot answer this all at once.

The balance, however, will tell us whether iron in rusting loses or gains anything, or whether the iron is simply changed into rust.[2]

EXPT. 1. **To find whether Iron changes in Weight when it rusts.** Weigh a watch-glass. Place upon it 3 or 4 grams of bright iron filings, and a long iron nail. Weigh again. Moisten the filings thoroughly with a little water, and

1 This inquiry was suggested by Prof. H. E. Armstrong, B. A. Report, 1889.

2 The teacher will probably find it advisable to carry out himself most of the experiments in this research before the class. They require time, and may suitably be continued while the next research on the Discovery of Oxygen is being pursued. For instance, Experiments 1 and 2 may be started together in the course (perhaps best at the end) of one day, and the results examined during the next.

stir thoroughly, carefully avoiding the loss of any filings. Place the watch-glass in a warm place near a fire or stove. Cover it with an inverted beaker or shade. Allow it to stand for one or two days. Then place the glass upon a sand-bath, and warm it until quite dry. Weigh. There is *a gain in weight*. This must have arisen either from the air or from the water or from both.

EXPT. 2. **To find whether some of the Air disappears when Iron rusts in it.**—Cut a cork so that it can be pushed up to the top of a glass jar. Bend a pin or a small piece of wire into a hook and fix it to the cork. Make a small bag of fine muslin. Place some bright iron filings in it, moisten the bag, and suspend it. Push the cork right up to the closed end of the jar. Place the jar mouth downwards in a basin or beaker containing a little water (Fig. 15). Put the whole aside in a

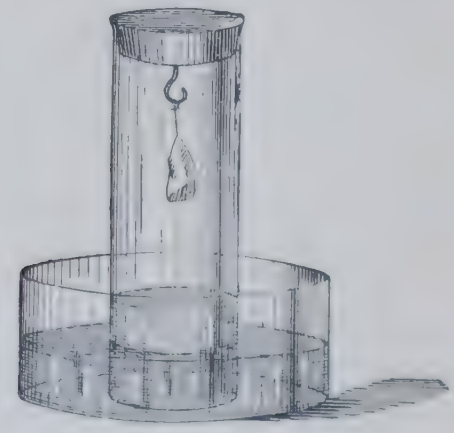

FIG. 15.

warm place until next lesson. After a day or two the level of the water is found to have risen in the jar. Pour water into the beaker so that the level of the water is the same inside and outside the jar. (Why is this done?) Mark the level of the water in the jar. Leave it for another day, and then mark the level again. Repeat this. At last the level of the water remains fixed. Find what fraction of the air in the jar has disappeared by measuring with a rule the length of the jar and the length of the column of air which is left. About one fifth of the air has disappeared. Leave the jar as it is for Experiment 5.

We have now proved that when iron rusts in the presence of air and water—

1. The iron gains in weight.

2. About one-fifth of the air disappears when the rusting occurs in a closed volume of air.

It seems to be certain, therefore, that the increase in weight of the iron is due (in part at any rate) to a portion of the air becoming fixed in or combined with the iron. It does not at present seem clear from these experiments whether the water has any share in the rusting. Common experience has certainly led people to imagine that water is more than anything else the cause of rusting. Popular opinion, however, is perhaps at fault, and must not be too readily trusted.

Will it not be well to try next whether (1) water in the absence of air, and (2) air in the absence of water can cause iron to rust?

EXPT. 3. **To find whether Iron rusts in Water free from Air.**[1]—Fill a small flask with distilled water. Boil it for at least five minutes, so as to expel all dissolved air. Meantime rub some iron nails bright with emery paper. Place them in the boiled water, and close the flask with a tightly-fitting rubber stopper. Set it aside till next lesson. No rusting occurs even after many days.

EXPT. 4. **To find whether Iron rusts in Air free from Moisture.** – Take a dry glass jar. Suspend a bag containing bright iron nails or filings within the jar exactly as in Experiment 2. Place the jar mouth downwards upon a well-greased plate, and set it aside in a warm place till next lesson (Fig. 16) Then try whether any of the air has disappeared, by removing the plate from the mouth of the jar under water. The water does not rise in the jar. Moreover, the iron remains bright.[2]

Experiments 3 and 4, therefore, satisfy us that neither water alone nor air alone can be the cause of the rusting of iron, Consequently, it is clear that iron only rusts in air when the air is moist. The formation of rust on iron is then due to two actions: firstly, the action of the air; and secondly, the action of moisture. Although in the present instance we shall only consider the first action in detail, the student should always

[1] Experiments 3 and 4 should be started together at the end of a lesson.

[2] Ordinary air always contains a little invisible aqueous vapour, and consequently the iron may rust to a very small extent.

accustom himself in studying an action to observe so far as possible all that goes on and all the conditions necessary for its success.

Now when iron was allowed to rust in moist air in Experiment 2, some of the air, but not the whole, disappeared and combined with the iron. Moreover, the fraction of the air which disappeared was *one-fifth*, which was exactly the same fraction of the air that vanished when phosphorus was *burnt* in the air (*see* p. 18, Experiment 8). Is it the fact then that that portion of the air which disappeared in Experiment 2 was simply the Fire-Air which it contained? If so, the air which remains must be Spoilt Air. We can put this to the test.

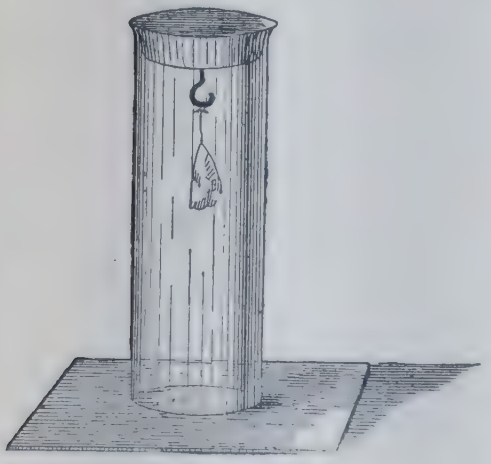

FIG. 16.

EXPT. 5. **To find whether the Air left after the Rusting of Iron in it is Spoilt Air.**—Fill the beaker containing the jar in Experiment 2 up to the brim. Gradually raise the jar until you can push a well-greased plate underneath it. Now invert the jar on the table with the plate uppermost. Light a candle. Also place a little red phosphorus in a deflagrating spoon, and kindle it. Remove the plate, and at once plunge the burning phosphorus within the jar. It is extinguished. Introduce quickly the burning candle. It also is put out.

We can safely conclude, therefore, that the air which remained after the rusting was Spoilt Air.

The Rusting of Iron is a Sort of Slow Burning. We have therefore proved beyond doubt that in the rusting of iron in a closed volume of moist air the observed gain in weight is due, in part at any rate, to Fire-Air combining with the metal. Now when iron burns in air the iron also combines with the Fire-Air; but in this case the action is completed in a few seconds, while the rusting of iron takes place very slowly. It seems, therefore, as though in the rusting of iron there is really

a very slow action between Fire-Air and iron, so slow that no heat nor light that we can detect is produced. It seems, indeed, to be a case of very slow combustion (*see* p. 25).

THE SMOULDERING OF PHOSPHORUS.—The above interesting result leads us to ask whether any other substance can combine without flame in this slow and gradual way with Fire-Air. Phosphorus is a likely substance to try for this slow combustion, since it begins to smoulder and fume as soon as it is exposed to the air.

EXPT. 6. **To observe the Smouldering of Phosphorus.**—Remove a small piece of yellow phosphorus from the bottle. Dry it quickly by pressing it gently between folds of filter-paper. Place it upon a porcelain crucible lid. A white smoke begins to arise from the phosphorus. Sometimes the phosphorus will soften and melt, showing that it is warm, and occasionally it will even become hot enough to take fire by itself.

Take the phosphorus to a dark place. It is seen to glow. Hence its name *phos-phorus*, the light-bringer, from the Greek words *phos*, light, and *phero*, I bring.

EXPT. 7.[1] **To find whether some of the Air disappears when Phosphorus is left exposed in it.**—Obtain a glass tube about 50 cm. long and 12 mm. broad, closed at one end. Fill it with water. Take a piece of copper wire about 40 cm. long, and twist one end into a small loop. Place a little water in a test-tube, and drop into it a piece of phosphorus about the size of a haricot bean. Melt the phosphorus by dipping the small test-tube into a beaker full of hot water, and then push the looped end of the wire into the melted phosphorus. Remove the small test-tube from the hot water, and dip it into cold water: the phosphorus solidifies and sticks to the wire. Push the end of the wire bearing the phosphorus half way down the tube filled with water, and bend the wire outside the tube back along the tube. Close the open end with the thumb, and invert it in a tumbler containing some water. Lift the open end of the tube for a moment out of the water, so that all the water runs out of it, and again lower it, and support the tube in a retort-stand (Fig. 17). The phosphorus is then exposed in a closed volume

1 Adapted from Parrish and Forsyth's *Chemistry*, p. 14.

of air. White fumes are seen, and in the course of some minutes the water is seen to have risen from the tumbler a little way up the tube. Set the tube aside till next lesson. The water is found not to rise beyond a certain point, and yet there is still plenty of phosphorus. Estimate what fraction of the air at first in the tube has disappeared. Is it not about one-fifth?

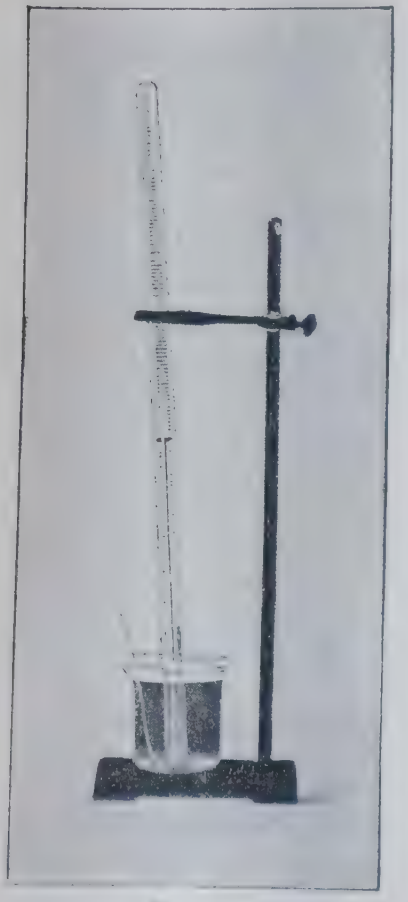

Fig. 17.

Then, taking care not to lift the end of the tube out of the water, pull the wire out with the phosphorus, and put it in some water to avoid accident. Close the open end of the tube with the thumb, and invert the tube with the water it contains. Test the air which remains in the tube with a burning splint: it is extinguished at once. Pour the water in the tube into a measuring jar, and so find the volume of the air which has disappeared. Also find the whole volume of the tube, *i.e.* the original volume of atmospheric air enclosed. Estimate again the fraction of the air which disappeared.

The Smouldering of Phosphorus is a Sort of Slow Burning.—We have now established the following points with regard to that which occurs when phosphorus is exposed in a closed volume of air.

1. One-fifth of the air disappears.
2. The air which is left has the properties of Spoilt Air.

We have therefore strong reasons to conclude that phosphorus combines even at the ordinary temperature with Fire-Air. The action appears to be precisely the same in kind as that which occurs when phosphorus is burnt in air, and the difference between the two cases seems to be simply one of rapidity and intensity.

Phosphorus, like iron, then, can combine with Fire-Air either very slowly or extremely rapidly. In the one case no action is visible, in the other much heat and a brilliant light are produced. It seems probable, indeed, that the heat and light in the second case are the *result* of the rapidity and violence of the action between the Fire-Air and the phosphorus or iron. This, it will be remembered, was the conclusion we came to at the end of our last research (*see* p. 25).

When the combination of phosphorus with Fire-Air is extended over hours or days, it seems probable that the action is scarcely visible because it is so slow. A faint glow can, however, be seen if the experiment is made in a dark room.

We have, therefore, confirmed by this study of the slow changes which phosphorus and iron undergo in the air the conclusion we had already come to about Flame and Fire. But this was not at all the actual object we had in view when we began to study the rusting of iron. It is not, however, any less valuable on that account, and as a matter of fact, it frequently happens that in the pursuit of knowledge we set out with some one object in view, and we arrive at an entirely different and unexpected discovery, yet perhaps even more important. But if our researches are to lead us thus into wider knowledge, we must be careful to keep our minds open, quick to observe any new fact, and willing to endeavour to understand it.

Problems

1. Observe the effect of heating iron rust in a test-tube.
2. Find whether there is any change of weight when iron rust is strongly heated in a crucible.
3. Find the action of hydrochloric acid upon iron rust. Is any Fixed Air evolved?
4. Repeat Experiment 2, p. 28, using strong sulphuric acid instead of water.
5. Find out what occurs when powdered calcium chloride is left on a watch-glass in the air.
6. Find out what occurs when a beaker half full of strong sulphuric acid is left exposed in the air.
7. Select some clear crystals of washing-soda. Leave them exposed to the air. What do you notice? Find out what has occurred.

CHAPTER IV

THE DISCOVERY OF OXYGEN

THE work of the Oxford chemists of the seventeenth century showed that common air was probably composed of two different gases, which we have called Fire-Air and Spoilt Air. No one, however, had succeeded in isolating Fire-Air. It is evident, therefore, that though these chemists made a great advance towards a knowledge of the air and of combustion, more remained to be done. As a matter of fact, no further advance was made until the latter part of the eighteenth century, through the investigations of *Scheele* in Sweden, of *Priestley* in England, and of *Lavoisier* in France.

THE SEARCH FOR FIRE-AIR

It must now be our object to isolate Fire-Air. We have found that, when metals and phosphorus are heated in the air, they increase in weight, because they form compounds with the Fire-Air. Is it not possible that we may be able to separate the Fire-Air again from some one of these compounds? For instance, when metallic lead is heated for a long time in the air, substances called *Red Lead* and *Litharge* may be obtained, which we are regarding as compounds of lead and Fire-Air. Now, is it not possible that by some treatment we may induce the lead and the Fire-Air to part company? An experiment in which this was effected was first made in 1774 by *Dr. Joseph Priestley*,[1] an eminent dissenting minister.

[1] Priestley says at the outset of his account of his experiments: "more is owing to what we call *chance*, that is to the observation of events arising from unknown causes, than to any proper design". He was ignorant of the work of the Oxford chemists, and after working on Fixed Air proceeded in a hundred and way to heat red lead and many other substances, simply to see what kind of air they would yield (see Alembic Club Reprint, No. 7).

This is how he carried it out. He wrote :—

"Having procured a lens I proceeded with great alacrity to examine by help of it, what kind of air a great variety of substances would yield, putting them into the vessels represented (Fig. 18), which I filled with quicksilver and kept inverted in a basin of the same."

Amongst other substances he tried **red lead**. He found that an air was expelled from it, and at the same time the red lead became yellow like litharge. He got sufficient of the air to try a few experiments with it, and he says: "What surprised me more than I can well express was that a candle burned in this air with a remarkably vigorous flame, and a piece of red-hot wood sparkled in it and consumed very fast."

FIG. 18.

We cannot readily repeat the experiment in just his way; for one reason, we cannot always have the sun's rays; but we can heat some red lead in a test-tube.

EXPT. 1. **A Fire-Air is obtained from Red Lead.**—Place some red lead in a small narrow test-tube. Heat it in the full flame of a Bunsen burner. It at first darkens in colour and then partly fuses. Introduce into the tube a glowing splinter of wood, and notice that it bursts into flame (Fig. 19).

It seems likely that the gas which the red lead evolves is the Fire-Air we are in search of. Note, too, that the residue is not a metal, but is like the earth first obtained on roasting lead, called litharge.

PROBLEMS

Find whether any of the following substances yields a Fire-Air when heated in a Bunsen flame: use a small test-tube, and take as much of each as would cover a sixpence.

Calcined mercury.
Litharge.
Calx of copper.
Calx of magnesium.
Calx of tin.
Nitre.

In the case of **nitre** compare what Scheele,[1] a poor Swedish

[1] It should be noted that Scheele discovered Fire-Air before Priestley, about the year 1772. Priestley was ignorant of this (*see* Scheele's *Fire and Air*, Alembic Club Reprint, No. 8).

apothecary, observed about the year 1772. He says: "I put an ounce of purified nitre in a glass retort for distillation, and made use of a bladder, moistened and emptied of air in place of a receiver. As soon as the nitre began to glow it also began to boil, and at the same time the bladder was expanded by the air that passed over. This is the cheapest and best method of obtaining Fire-Air."

In the course of his experiments Priestley heated the red substance called **calcined mercury**. It is formed upon the sur-

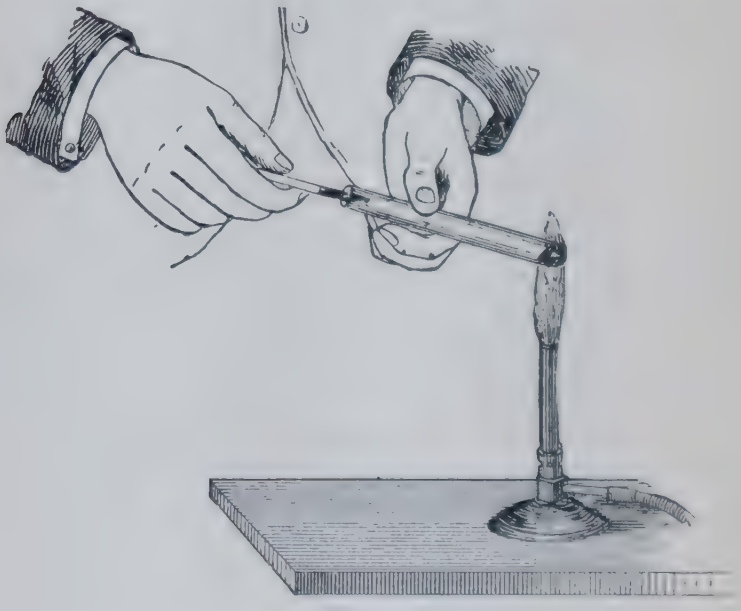

FIG. 19.

face of the mercury when mercury is heated in common air. He found that an air was readily given off.

EXPT. 2. **A Fire-Air is obtained from Calcined Mercury.** Place a small quantity of calcined mercury[1] in a test-tube. Heat it cautiously but strongly in a Bunsen flame. It darkens in colour. Introduce a glowing splinter of wood (Fig. 19). The wood bursts into flame. Observe also that a mirror of minute globules of mercury is deposited upon the cooler portion of the tube, which on being rubbed gather together into a drop of the liquid metal. Thus calcined mercury yields mercury and a Fire-Air.

[1] Known to the chemist as oxide of mercury.

Note that it has not yet been proved that the Fire-Air obtained from calcined mercury is the *same* as that present in common air.

As we have said, Priestley discovered this Fire-Air in 1774; in the same year he told Lavoisier in Paris about it, and the famous French chemist at once repeated the experiment in a very instructive way.

Lavoisier's Experiments.—(1) Lavoisier took a glass retort, A, with a long neck, and after bending the neck as repre-

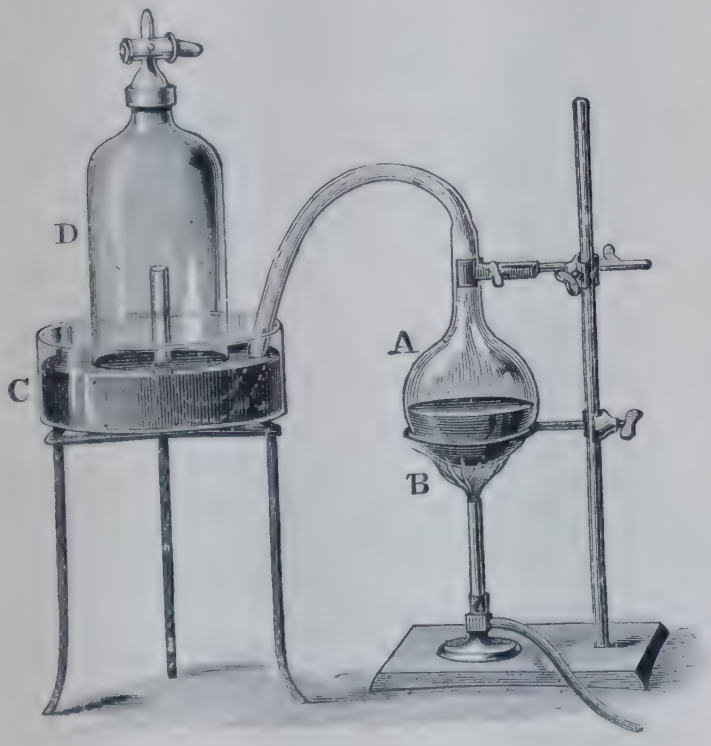

FIG. 20.

sented in Fig. 20, he introduced some mercury, and placed the retort over a flame. The end of the neck came out above the surface of mercury in a dish, C, and under a graduated bell-jar, D, containing 50 cubic inches of air. The mercury in the retort was heated strongly by the flame for several days. As this went on, red specks of calcined mercury formed on the surface in the retort, while the mercury gradually rose in the jar. After a few days, *no further formation* of the red substance appeared to

be taking place, so after twelve days from the outset the experiment was stopped. It was found that the volume of the air in the jar had decreased between 7 and 8 cubic inches, and that which remained in D was not fit for respiration nor did it support combustion. Then Lavoisier collected the red calcined mercury. It weighed 45 grains.

(2) The calcined mercury was next placed in a glass tube closed at one end, and bent as represented in Fig. 21, the open

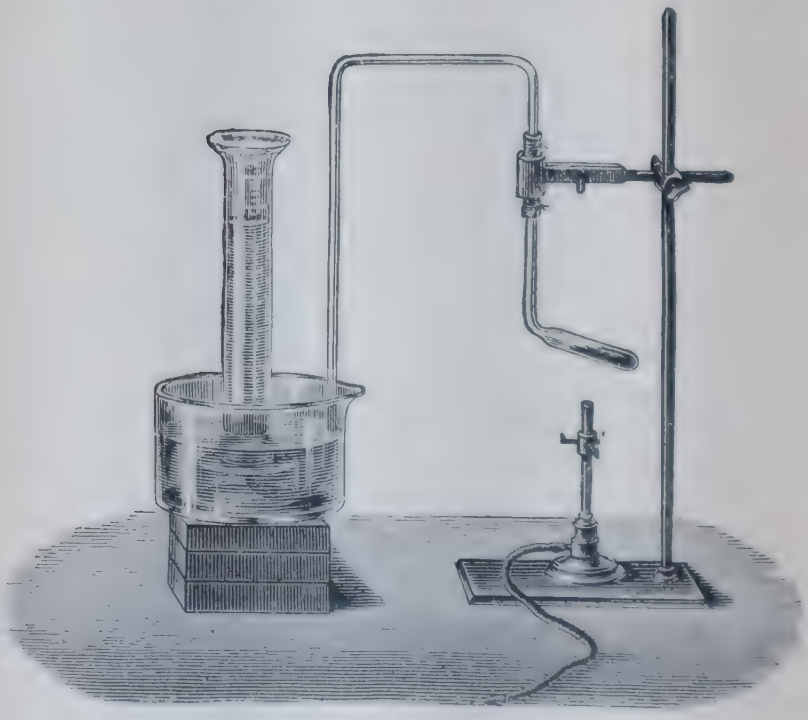

FIG. 21.

end being placed under a graduated cylinder filled with mercury. On heating the substance, a gas was given off, which measured between 7 and 8 cubic inches, and 41½ grains of mercury were deposited on the colder parts of the tube. The gas collected in the cylinder was proved to support vigorously the combustion of a taper or charcoal.

(3) Lavoisier repeated the analysis of 50 cubic inches of air as described above, and then added the 8 cubic inches of gas obtained on heating the calcined mercury as described under (2) to the 42 cubic inches left in the bell jar. He *obtained*

in this way an air which was in every respect like atmospheric air.

Therefore, after having by **analysis** proved common air to contain in reality two airs or gases, he confirmed this by the method of **synthesis,**[1] that is, by putting the two airs together **and so producing common air again.**[2]

Let it be particularly observed that the volume of gas obtained from the calcined mercury by heat was the *same* as the volume of gas which disappeared from the bell-jar in Lavoisier's Experiment (1), p. 37.

The experiments make it quite clear that the gas which calcined mercury evolves on being strongly heated is the very *same* which mercury picks out of the air, when calcined mercury is made. Therefore, **the gas which calcined mercury gives off on being heated is the same as that portion of the air which we have called Fire-Air.** We have thus found a method of preparing Fire-Air, and we can now proceed to try experiments in order to find out its characteristic properties.

Calcined mercury is an expensive substance, and we cannot therefore afford to use it to obtain sufficient Fire-Air in order to try many experiments. After it had been discovered that Fire-Air could be obtained from calcined mercury, other methods of preparing it were very soon found, and we will make use of one of them. Berthollet, about 1800, found that chlorate of potash evolves Fire-Air when heated. Let us try also the effect of heating it when mixed with other substances.

EXPT. 3. **Preparation of Fire-Air from Chlorate of Potash.**—Heat the following substances in a dry test-tube in a Bunsen flame, and try to compare the ease with which Fire-Air is evolved, detecting it by means of a glowing splint.

(1) Chlorate of potash (powdered).
(2) Pyrolusite.[3]
(3) Chlorate of potash mixed with a little pyrolusite.
(4) Chlorate of potash mixed with a little sand.

The results suggest the following method of obtaining readily a steady supply of Fire-Air.

[1] Greek: *syn*, together, and *thesis*, a placing.

[2] Lavoisier's Experiment No. 1 can be illustrated by heating some mercury for some days in a round-bottomed flask on a sand bath, the flask being connected, by a glass delivery tube, bent four times at right angles, with some air in a test-tube standing over mercury.

[3] Also called *manganese dioxide*.

EXPT. 4. **Laboratory Preparation of Fire-Air.**—Take a large test-tube, 6 × 1 inch, and fit it with a cork and delivery tube, 50 cm. long, bent as shown in Fig. 22. Weigh out 20 gm. of chlorate of potash and 5 gm. pyrolusite. Dry each by warming them gently in porcelain basins, and stirring with a glass rod. Grind them together in a mortar: add about an equal volume of sand, and mix well. Place the mixture in the test-tube, and fix the tube as shown, in the clamp of a retort-stand. Arrange to collect the Fire-Air over water. Have six jars

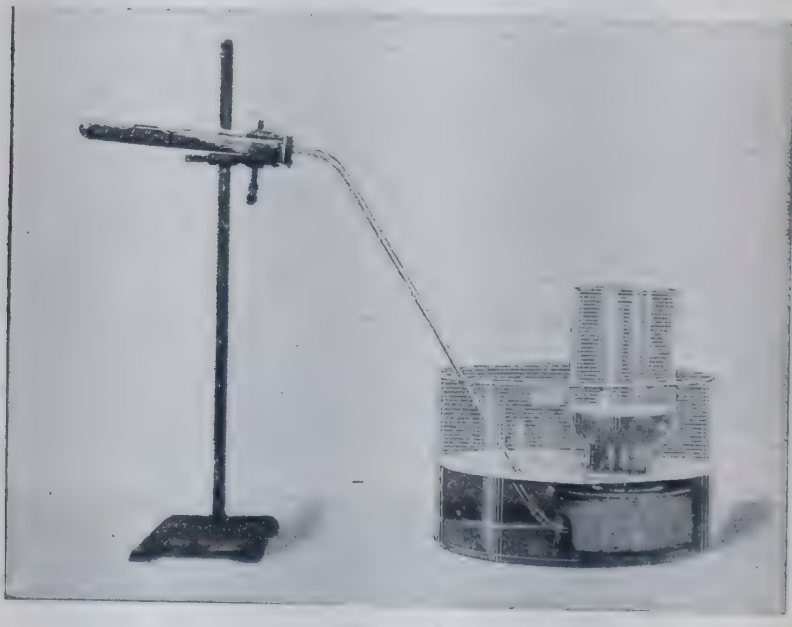

FIG. 22.[1]

ready. Heat the tube cautiously by moving a Bunsen flame to and fro along the tube. After collecting six jars of Fire-Air, remove the delivery tube from the trough. Let the test-tube cool, and then wash it out.[2]

FIRE-AIR VIGOROUSLY SUPPORTS COMBUSTION. We may now proceed to test the characteristic properties of Fire-Air. Remember that it forms one-fifth by volume of common air, and the chief difference between it and common air is that it is not mixed with Spoilt Air. What difference is this likely to make between the burning of phosphorus, sulphur, a candle, or iron, in it and in air?

[1] From Parrish and Forsyth's *Chemistry*, p. 53.

[2] The gas is also very conveniently obtained compressed in steel cylinders.

EXPT. 5. **To burn a Candle in Fire-Air.**—Attach a small candle to a deflagrating spoon. Light it at a flame, and introduce into a jar. It burns with a more brilliant white flame than in common air, and the wax melts very rapidly.

EXPT. 6. **To observe the Burning of Sulphur in Fire-Air.**—Place a piece of sulphur in a deflagrating spoon. Pour a little water into a gas jar. Direct a flame on to the sulphur until it catches fire, and introduce it into the jar. The faint blue flame at once burns more brilliantly, and white fumes are formed. When the flame dies out, remove the spoon, and observe the suffocating odour of the fumes. Cover the mouth of the jar with the hand, and shake. The fumes are dissolved in the water. Add some litmus solution. It is turned red, showing that an acid has been formed. We may call it *sulphurous acid.*

EXPT. 7. **To observe the Burning of Phosphorus in Fire-Air.**—Place a little red phosphorus in a deflagrating spoon. Kindle it in a flame, and introduce it into a jar. Observe that it burns with a very brilliant white light, and dense white fumes are produced, which readily dissolve in a little water. Add some litmus solution. It is turned red, showing the presence of an acid. We may call this *phosphoric acid.*

EXPT. 8. **To observe the Burning of Iron in Fire-Air.**—Fix a piece of a thin watch-spring or fine clean iron wire to a deflagrating spoon. Dip the end of the spring in some burning sulphur. Introduce it into another jar in which some water (about an inch in depth) has been placed. The metal becomes red-hot and then burns brilliantly, emitting sparks, and fused globules fall into the water. Examine these. They consist of a fused black calx of iron.

These experiments make it abundantly clear that those substances which burn in air burn much more brilliantly in Fire-Air. As Priestley expressed it, there is a "*superior goodness about this kind of air.*" Some words Scheele wrote in 1777 also sum up the matter very well: "These experiments show, therefore, that this Fire-Air is just that air by means of which fire burns in common air; only it is there mixed with a kind of air which places some hindrance in the way of the otherwise rapid and violent inflammation."

Fire-Air is Respirable.—We have previously come to the conclusion that Fire-Air is the only portion of the air which

supports respiration and the life of animals (p. 23). How then will pure Fire-Air suit them? Both Priestley and Scheele made experiments with animals, and found that they were much more lively and lived more vigorously in Fire-Air than in common air. Here is an interesting passage Priestley wrote in 1775:—

"My reader will not wonder, that, after having ascertained the superior goodness of Fire-Air[1] by mice living in it, and the other tests above mentioned, I should have the curiosity to taste it myself. I have gratified that curiosity by breathing it. The feeling of it to my lungs was not sensibly different from that of common air; but I fancied that my breast felt peculiarly light and easy for some time afterwards. Who can tell but that, in time, this pure air may become a fashionable article in luxury. Hitherto only two mice and myself have had the privilege of breathing it!"

Lavoisier observed the same properties of the new gas, and he often called it "*vital air*" or "*eminently respirable air*."

It may be remarked that Priestley's prediction has in a measure been fulfilled, for at the present time Fire-Air is in use in our hospitals, when it is desired to help the breathing of patients.

Lavoisier's Theory of Combustion.—In the year 1774, when Scheele and Priestley had independently discovered Fire-Air, Lavoisier was aware that sulphur and phosphorus, or the metals tin and lead, increased in weight when heated in the air. He put this down to common air getting "fixed" in them; but until he heard from Priestley of his discovery he did not thoroughly understand what it was that occurred in combustion. Then when he heard of the preparation of "Fire-Air," as Scheele called it, he made those experiments which we have described, and he saw clearly that in any case of combustion in the air whether of phosphorus, or of a metal, or of a candle, or of anything else, there was a **union taking place between it and that portion of the common air which we have called Fire-Air.**

This was a wide statement to make, and one which applied to a great many cases. The Oxford chemists had 100 years before come to very much the same conclusion (*see* p. 17), but every one had forgotten their work. At the time of the dis-

[1] Priestley gives it another name.

covery of Fire-Air almost every one believed in the Phlogistic Theory of Combustion, and thought that when a substance was burnt a fiery principle *escaped* from it. In direct opposition to this, Lavoisier maintained that in combustion the Fire-Air of the atmosphere *combined* with the substance burnt, so that the substance produced in the combustion was more complex, not simpler, than that which had been burnt. As a matter of fact Lavoisier's experiments and his clear statement of the case very soon convinced almost every one of the truth of his theory of combustion, and the former Phlogistic theory was abandoned.

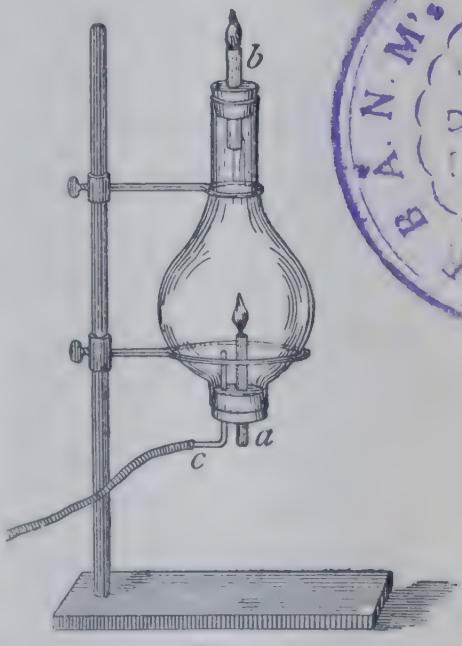

FIG. 23.

Combustibles and Supporters of Combustion.—When a piece of phosphorus is burnt in Fire-Air, the phosphorus is said in common language to be the *combustible* substance, and the Fire-Air *the supporter of the combustion.* According to Lavoisier's theory, however, the combustion consists in the union of the two, and it is as correct to say that the phosphorus combines with the Fire-Air, as that the Fire-Air combines with the phosphorus. Take the case of coal-gas burning in air; it is usually said that the air is the supporter of the combustion, while the coal-gas is the combustible substance. Here we have two gases. If the combustion consists only in the union of the two, why should we call one the combustible substance any more than the other? The reason is that usually coal-gas is burnt *in* air. But *can* we burn air in coal-gas? Let us try.

EXPT. 9.[1] **To find whether Air will burn in Coal-Gas.** Obtain a lamp glass. Fit the wider end with a cork in which are fixed a straight piece of wide glass tubing, *a* (Fig. 23), about 10 cm. long, and also a narrower piece of glass

[1] Lecture Table Experiment.

tubing, *c*, bent at right-angles. Fit the upper end also with a cork in which is fixed a short glass tube, *b*. Support the lamp glass in the ring of a retort-stand. Connecting the tube *c* with the coal-gas supply, pass coal-gas into the glass, and light the gas as it issues from *b*.

Then, without turning off the gas, close the end of the tube *b* with a cloth, apply a flame to *a*, and at once remove the cloth. Notice that the flame which a moment before was burning at the end *a*, has travelled up the tube and is now burning at the other end. This flame, moreover, must be one of *air burning in coal-gas.*

It is seen, therefore, that the terms "combustible" and "supporter of combustion" are not used in a strictly accurate sense. They are, however, convenient.

The fact that air can be burnt in coal-gas just as coal-gas in air, supports Lavoisier's theory of combustion. Could it be explained by the former Phlogistic Theory?

THE COMPOSITION OF FIXED AIR.—The student has already observed that a piece of charred wood burns away and leaves hardly any ash. Since charcoal burns in air, it will probably burn all the more vigorously in pure Fire-Air. Let us try this.

EXPT. 10. **To observe the Burning of Charcoal in Fire-Air.**—Take a large flask or cylinder and fill it with Fire-Air. Place a piece of charcoal (charred bark best of all) in a clean deflagrating spoon. Heat the charcoal in a flame until it is kindled, and then plunge it into the Fire-Air. It burns with great brilliancy. No fumes are seen, yet the piece of charcoal is smaller. The only visible product is the white ash which now coats what is left of the charcoal.

What has become of the charcoal? Have we any clue? Yes, Lavoisier's Theory of Combustion. This theory suggests the idea that the charcoal (or *carbon* as it is often called) has united with the Fire-Air, and formed some invisible air. With what different airs are we already acquainted? They are common air, Fire-Air, Spoilt Air, and Fixed Air. To see whether perchance any Fixed Air is formed, pour into the flask about 20 c.c. of clear lime-water, cover the mouth with the palm of the hand, and shake well. A turbidity is at once seen. This proves

that *Fixed Air has been produced by burning carbon in Fire-Air.*

It might be suggested that the Fixed Air has been driven out from the carbon ; but, in view of Lavoisier's Theory of Combustion, it is more probable that the Fixed Air has been formed as a result of the *union* of the carbon with Fire-Air. Let us, at any rate, proceed on this supposition as a " working hypothesis," and try whether we can separate carbon from Fixed Air. How is it likely we can do this? What substances can combine with Fire-Air, which we think there is in Fixed Air? Several metals do so, and magnesium especially readily. Let us try.

EXPT. 11. **Does Fixed Air contain Carbon?**—Fill a jar with Fixed Air. Fix a strip of magnesium ribbon to a deflagrating spoon. Kindle the end of the ribbon, and plunge it into the jar. The metal continues to burn. A white substance is formed like that produced when magnesium is burnt in air. but it is mixed with black specks, which look like charcoal or carbon. Add some strong hydrochloric acid, and shake. The calx of magnesium is dissolved, and black specks are readily visible. Filter. Devise some means of proving that these specks are really carbon, and not some other black substance. Will it mark paper, and glow when heated on platinum foil? Yes.

We have, therefore, confirmed by analysis our conclusion (arrived at by synthesis, Experiment 10) that **Fixed Air is a compound of carbon and Fire-Air.** Notice that it was the application of Lavoisier's theory of combustion which led up to this discovery.

A new Name for Fixed Air.—When we were first studying Fixed Air it was noticed that it possessed an acid character, for it was found to turn litmus solution red (p. 8). To express the fact that Fixed Air is an acid, and that it is a compound of carbon, Lavoisier proposed to call it **Carbonic Acid** a name which has been in use ever since, and one which we shall now employ.

A new Name for Fire-Air.—We have found that when carbon is burnt in Fire-Air, carbonic acid is produced. The student will remember also that when sulphur and phosphorus are similarly burnt acids are also produced, which we have called sulphurous acid and phosphoric acid. These facts were

observed by Lavoisier, and led him to regard Fire-Air as a gas which had the property of producing acids when it combined with many substances. He therefore proposed to give Fire-Air the name **oxygen**—a word which means the *acid-producer*. It is derived from two Greek words *oxus*, acid, and *gennao*, I produce. This name gradually supplanted the old one, and is employed by every chemist at the present day for this gas.

At about the same time the name **nitrogen** was given to Spoilt Air, and it is the name now employed.[1]

The Definition of an "Element."—Many of the substances with which we have been dealing can no longer be regarded as simple in nature. The air around us contains two different gases, oxygen and nitrogen. Calcined mercury when heated liberates oxygen, and mercury is left. Red lead, again, is a compound of lead and oxygen. On the other hand, we have no evidence whatever that oxygen, or mercury, or lead are compound in nature, and we seem to require some word to distinguish such substances from those which are known to be more complex.

Such a word is at hand in the term **Element**, one which is in frequent use in ordinary language, and one which has for hundreds of years been employed by alchemists and chemists. But the Honourable Robert Boyle was the first to state clearly the sense in which it should be used. In his book *The Sceptical Chemist*, he pointed out in 1661 that only substances which can be obtained from compounds, and are not known themselves to be compounds should be regarded as elements.

DEFINITION.—**An Element is a substance which has not by any known process been separated into two or more different substances.**

Then, lead, iron, and oxygen are, so far as our knowledge goes, examples of elements. Carbonic acid gas, on the other hand, is not an element, for we have been able to separate carbon from it, and we know moreover that it contains oxygen.

The Definition of an Oxide.—We have been led to recognise the presence of oxygen in the red calx of lead, in the

[1] This name was suggested by Chaptal in reference to the fact that Spoilt Air could be obtained from *nitre*. Nitrogen was discovered by Rutherford of Edinburgh in 1772. Lavoisier called it *azote* (*a*, not, and *zoe*, life).

white substance formed from burning phosphorus, and in the pungent smelling gas which arises from burning sulphur, as well as in carbonic acid gas. All these substances can be looked upon from the same point of view as compounds of oxygen with another substance, which, so far as our knowledge goes, is an elementary one. Thus, though they differ greatly from one another in appearance and in properties, yet from a chemical point of view they are all constituted in a similar way.

It is convenient to call any such substance an *oxide*.

DEFINITION.—**An Oxide is a compound of an element with oxygen.**

A name which is often used for carbonic acid is *carbon dioxide*. The full significance of this cannot be understood at present, but it refers to the fact that the gas is a compound of carbon and oxygen.

TABLE OF SYNONYMS OF OXIDES

It will be useful to tabulate for purposes of reference the old names and also the names which have been commonly adopted since the work of Lavoisier for some of the chief oxides :—

Litharge	Lead monoxide
Rust of iron	Ferric oxide
Calx of tin	Stannic oxide
Calx of zinc	Zinc oxide
Calcined mercury	Mercuric oxide
Calx of magnesium	Magnesium oxide, magnesia
Pyrolusite	Manganese dioxide
Fumes of burning phosphorus	Phosphorus pentoxide
Fixed Air	Carbon dioxide, carbonic acid
Fumes of burning sulphur	Sulphur dioxide

Matter is indestructible.—The thoughtful student has perhaps noticed that on more than one occasion in the preceding pages we have *assumed* that matter[1] is indestructible, and that it cannot be created. For instance, when we burnt charcoal in oxygen (Expt. 10, p. 44), we saw the charcoal disappear, but we took it for granted that it had not been annihilated. We argued that it had combined with the oxygen forming some invisible gas, and we proceeded to try to find what this was, and

[1] DEFINITION.—*Matter* (or substance) is that which has (1) weight, (2) inertia, (3) size.

found it to be carbonic acid (Fixed Air). Then again, when we heated lead in the air (p. 15), and observed an increase in weight, we assumed that this could only be accounted for by some portion of the air becoming attached to the metal; in other words, we did not regard the possibility of more lead having been created, but attributed the gain in weight to the addition of some external matter.

Now the first impression of a child observing a candle disappear, or a small seed growing into a big plant, is that matter can be utterly destroyed, or, as in the case of the seed, grow and increase "of itself." Let us consider therefore what grounds we have for assuming, as we have done, that matter is everlasting and indestructible.

1. *Scheele's experiment* (p. 18), in which a piece of phosphorus was burnt in a closed flask full of air, verified this in a very clear way. For no change in weight had occurred after the combustion. This showed that the heat applied had no effect either in increasing or decreasing the weight of the original substances, and that no loss of weighable matter had occurred, although the phosphorus had burnt in part away. Therefore, when the flask was opened under water, and it was found that one-fifth of the air had disappeared, we were perfectly justified in saying that this air had not been destroyed, but had combined with the phosphorus, forming the white solid substance. Thus, there were present in the flask:—

Before the combustion	*After the combustion*
Phosphorus	Unburnt phosphorus
Oxygen (Fire-Air)	Oxide of phosphorus
Nitrogen (Spoilt Air)	Nitrogen

In this case, therefore, the total weight of the substances after the chemical action was the same as the total weight at the outset.

2. Another illustration of the indestructibility of matter may be taken from *Lavoisier's experiments on the calcination of mercury* (p. 37). It was shown that when mercury was calcined in the presence of a given volume of air, the volume of the air was diminished by nearly one fifth. This air was not destroyed. It had simply combined with the mercury forming a red substance.

For when the calx of mercury was collected and heated it evolved exactly the same volume of air.

3. Again, a *burning candle* is apparently destroyed. But is this so? The difficulty is to kindle a candle in a *closed* volume of air.

EXPT. 12.[1] **To find whether there is a Loss of Matter when a Candle is Burnt.**—Take a round-bottomed flask of half a litre to one litre capacity. Select a tightly fitting rubber stopper. Take a wax vesta (which may be considered a small candle), and bind one end of a piece of thin copper wire firmly round the tail of the match. The wire should be about three-quarters of the length of the flask. Make a loop at the other end of the wire, and tie to it a piece of wool a few inches in length. Suspend the match within the flask so that its head is about half an inch above the bottom, and then fix the stopper in the neck, pressing the wool tightly against the glass. Make a loop in the wool outside the flask, and suspend the whole from one arm of a balance, and exactly counterpoise it. Then remove the flask from the balance, and incline it so that the head of the match touches the glass wall. Cautiously warm the flask in a Bunsen flame until the head of the match catches fire, and at once hold the flask upright, so that the match hangs freely. The match continues to burn for a short time, rather smokily, and then the flame gradually dies down and goes out.

FIG. 24.

A portion of the match has undoubtedly "burnt away." Has there been any loss of weight? Allow it to cool for at least ten minutes. Place the flask on the pan of the balance again with the same counterpoise. There is no change of weight. Then there has been no destruction of the material of the match.

These three instances will suffice to illustrate a law which is of the very highest importance in the study of chemistry, when the changes substances undergo are studied. This **Law of**

[1] Lecture Table Experiment, adapted from an experiment suggested by J. B. Coleman, in *The School World*, Feb., 1900.

the Indestructibility of Matter, or **Law of Conservation of Mass,** as it is often called, may be very briefly stated in the words :—

MATTER CANNOT BE CREATED AND CANNOT BE DESTROYED.

"We may lay it down," said Lavoisier, "as an incontestable axiom that in all the operations of art and nature, nothing is created ; an equal quantity of matter exists both before and after the experiment. Upon this principle the whole art of performing chemical experiments depends."

It has been found that heat has no effect either in increasing or in decreasing the weights of the original substances. *Heat* is clearly itself unweighable, and it must therefore be fundamentally different from *matter*, which is weighable. This distinction, which we owe to Lavoisier, between ponderable matter and an imponderable agent, such as heat, is one of the greatest discoveries that has ever been made.

The discovery of carbonic acid (Fixed Air) did much to lead to a belief in the indestructibility of matter. It is evident that until it was found that carbonic acid is produced when charcoal is burnt, it was natural to suppose that the substance of the charcoal was actually destroyed. So, again, the Law of the Indestructibility of Matter could not be grasped until the discovery of oxygen was made, and the part it played in combustion had been clearly set forth by Lavoisier.

Very truly did Scheele remark in 1777, "If the chemists of the preceding century had thought worthy of a more particular examination the elastic fluids resembling air which manifest themselves in so many operations, how advanced should we now be ! They desired to see everything in corporeal form, and to collect everything as drops in the receiver. This is now for the first time better inquired into, and the air has begun to be carefully examined ; and who is there that does not perceive the advantage which the results of such experiments carry with them ?"

Exercises and Problems

The Discovery of Oxygen

1. Find whether any of the substances, A, B, C, D, E, F,[1] yield Fire-Air (Oxygen) when heated in a Bunsen flame.

2. Find which of the gas jars G, H, I, K, contain common air, Fire-Air, Spoilt Air, Fixed Air, respectively.

3. Try to find what becomes of a candle as it burns away. Fix a small candle in a deflagrating spoon, light it, and introduce into a dry clean jar.

4. Find the effect of heating bright copper turnings in a hard glass tube in a current of oxygen.

5. Mix air and carbonic acid in gas jars in different proportions, and find whether a candle can burn in them all.

[1] *E.g.* HgO, BaO_2, PbO_2, $K_2Cr_2O_7$, Fe_2O_3, $KClO_3$

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CHAPTER V

THE DISCOVERY OF THE METALS[1]

THE METALS DISCOVERED BY THE ANCIENTS

MENTION has already been made in Vol. I. Cp. I. of the oldest extant chemical manuscript. Thanks, however, to the labours of archæologists we can state that chemical processes were practised in Egypt in times far more remote than those of the fraudulent goldsmith.

In many parts of the **British Isles** there are mounds known as barrows or tumuli, which are the sepulchral monuments of an early age. They often contain the remains of bodies, along with pottery, ornaments, and weapons or implements. These ornaments and weapons are usually of stone, but in some cases of bronze or iron, proving that the men, long since dead, who used them had some knowledge of the metals. Unfortunately it is usually impossible to assign any date to these tumuli.

GOLD AND COPPER.[2]—The case is different in **Egypt** and **Assyria**, where an advanced stage of civilisation had been reached when the British Isles were still steeped in barbarism. The ruins of ancient Egyptian cities have been explored, inscriptions and drawings on papyri, or clay tablets, or on the walls

[1] *To Teachers.*—This chapter may be postponed if desired. But, besides leading to an acquaintance with the sources of the common metals, it illustrates the gradual development of chemical knowledge and practice and ideas, and gives a training in accurate observation and deft manipulation.

[2] The Metals Discovered in Modern Times will be discussed in Vol. III. Much of the information in this chapter has been gathered from a Friday evening discourse delivered at the Royal Institution on Feb. 11, 1892, by Dr. J. H. Gladstone, F.R.S. (See *Nature*, lvii, 594.)

of palaces, tombs, and temples have been deciphered, rendering it possible to assign approximate dates to the numerous metallic objects which have been unearthed.

At Nagada, in Egypt, an enormous royal tomb was explored in 1897, which it is believed contained the mummy of the first Pharaoh, who died about 4400 B.C. Among many articles which it contained, of ivory, mother-of-pearl, tortoiseshell, and glass, there were some small pieces of metal—namely, a long bead of *gold*, and a bead and some fine wire of *copper*. These are the oldest metallic objects in the world to which we can give a probable date.

A few small copper implements have also been found at Nagada, which it is believed belonged to the inhabitants of the Nile Valley *before* the coming of the Egyptians; if so, they are very old.

The importance of these discoveries is that we may conclude that **the Egyptians at the very beginning of the historic period were acquainted with the use of the two metals, gold and copper.** There is no evidence that they knew of any other metal in these early years. It is interesting to observe that the working of gold is frequently represented in the Egyptian sculptures and pictures: goldsmiths may be seen making jewellery, weighing out the metal, melting it in their little furnaces with the aid of blow-pipe and pincers, washing it, and working it into the proper forms.

Why were gold and copper the first metals to be known? The answer is in part to be found in the fact that they both occur "native" in the earth.[1]

Gold occurs native, especially in quartz veins, and on account of its yellow colour readily attracts the attention. It was in early times found in great quantities in Asia Minor, in the sands of the Pactolus and of other rivers. The great Egyptian Kings Seti I. and Rameses II. (B.C. 1300) worked large gold mines in Abyssinia. At the present day it is obtained chiefly from Australia, California, the Transvaal, and Alaska.

Copper[2] is also found in the metallic state in fibrous masses,

[1] A metal is said to be "native" if it occurs as such, to be seen by the eye, in the crust of the earth. A mineral which contains a metal compounded with other material in which the metal is not recognisable, is called an "ore."

[2] Our word *copper* is derived from the late Latin *cuprum* and the Greek *kupros*, which record the fact that the Romans and Greeks obtained the metal chiefly from the island of Cyprus.

which may be easily hammered into any desired shape. We can account for the early knowledge of copper, also, in the fact that it is easily worked or smelted from some of its ores, for example, from the deep blue *azurite*, which occurs in Armenia, or from the dark-red *cuprite*.

In order that we may in the laboratory obtain the metals from their ores, we must learn the use of the blow-pipe.

To Use a Mouth Blow-pipe.—A blow-pipe (Fig. 25) will turn a sharply-pointed flame in any required direction. Close the air-holes of a Bunsen burner, and turn down the yellow flame, so that it is only one and a half inches in height. Point the blow-pipe jet a little downwards, and place the nozzle just within the flame, a little above and to the side of the burner, as

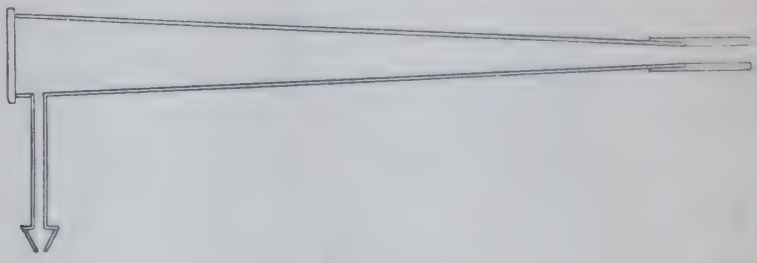

FIG. 25.

shown in Fig. 26. Blow gently, so that the flame just fails to hiss. A pointed blue inner flame results, surrounded by an almost invisible outer flame (Fig. 27). It is useful to be able to keep up a steady current of air for three or four minutes. To do this air must be breathed in through the nose by moving the tongue in the roof of the mouth; the cheeks will then be kept distended, and form a reservoir or supply of compressed air. Blow from the cheeks, not from the lungs. A metal mouth-piece is apt to make the lips itch; a short glass tube attached by a rubber connection will be more comfortable. Since air is added through the blow-pipe, the supply of air through the holes at the bottom is, as has been stated, cut off. The addition of air to the yellow flame changes it to a hotter blue flame; but if too much air is added the flame is cooled down again by the excess of air; when the flame is buzzing it shows that there is too much air.

Make a slight hollow upon a lump of charcoal, and place in

it a few particles of the metal *antimony*. Direct the tip of the inner blue blow-pipe flame upon the metal: it is rapidly melted. Now hold the metal in the tip of the outer flame. A dense white fume is evolved, and a white incrustation forms

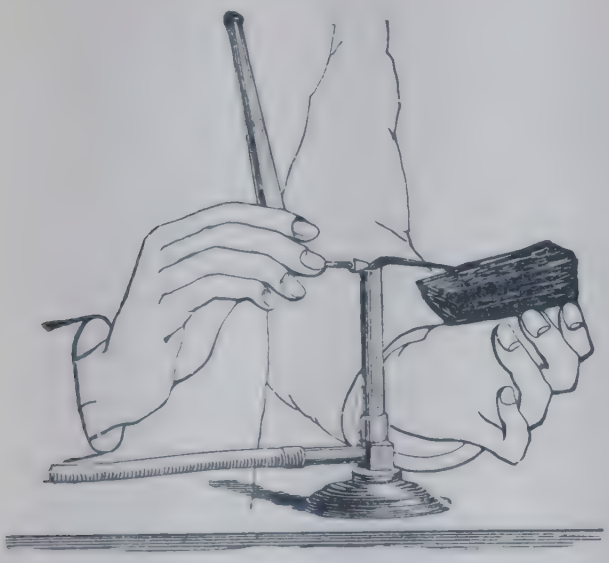

FIG. 26.

upon the charcoal. What is it? The oxide of antimony. (Cf. Vol. II. p. 16.) The outer flame is, of course, surrounded by intensely heated air, and it is the oxygen in this which is readily attracted by the metal when the outer flame is directed

FIG. 27.

upon it. Thus the outer flame is an *oxydising flame*. The character of the inner flame will become clear later.

The Use of a Flux. — In metallurgical operations it is usual to mix with the ore a material which will itself readily melt, and therefore allow any metallic globules contained in it to flow together. Such a substance is called a flux (Latin, *fluo*, I flow). The white substance[1] made by driving all the water out of

[1] Called "anhydrous sodium carbonate."

crystals of washing-soda will serve. There is a better, viz., the substance usually labelled "fusion mixture,"[1] which is preferred on account of its lower melting-point.

The Use of Charcoal.—In the laboratory it is usual to smelt ores upon lumps of charcoal. Charcoal is used partly because it is a non-conductor of heat, and may be made very hot at one place; but also because it is porous, and will absorb the flux and the impurities which the flux removes from the metal, so that the metal is left upon its surface. And further, charcoal is used because at the temperature of the blow-pipe flame it has an action upon many ores, tending to remove from them certain elements, particularly oxygen, with which the metals may be combined. (Cf. p. 44.)

EXPT. 1. **To obtain Copper from Azurite.**[2]—Grind a small quantity to a powder. Transfer a very small amount on the point of a knife to the palm of the hand. Add about four times as much fusion mixture. Slightly moisten the blade of the knife with water, and mix the materials with it into a thick paste, and transfer this to a slight hollow on a lump of charcoal. Direct the tip of the flame on to the mixture (*see* Fig. 26). The fusion mixture soon melts: then blow more strongly, and hold the charcoal so that the tip of the inner blue flame plays upon the mixture. Keep up a steady supply of air, and maintain the mass at a red-heat for two or three minutes. When cool, one or two little rough pieces of red metal can probably be seen. Transfer the metal with as little charcoal as possible to a small mortar. Crush it well. Hold the mortar so that a thin stream of water from a tap falls into it and carries off the charcoal. Examine the bottom of the crucible: several bright plates of copper can be seen. Notice that even at the high temperature of the blow-pipe flame the metal did not melt.

Try whether you can also obtain copper from *cuprite*.

TIN.—The oldest known copper articles consist of almost pure metal. But such copper is too soft to make efficient knives, razors, or carpenters' tools. There are, however, ways of hardening copper: for example, the addition of a small proportion of tin to copper forms the alloy or mixture of metals known as

[1] This is a mixture of the carbonates of potassium and sodium.

[2] Azurite is a carbonate of copper. The ordinary carbonate of copper may be used in this experiment if the ore is not available.

bronze, which is much stronger and harder than pure copper. It is certain, too, that Egyptians of an early age were acquainted with this fact. Bronze can be traced back as far as 3400 B.C., and at a somewhat later date bronze articles—knives, spear-heads, axes, statuettes, bracelets, and earrings—became common in Egypt.

There is reason, however, to doubt whether the Egyptians, at the early date of which we have spoken, were acquainted with tin, or were accustomed to harden copper by adding to it tin which they had separated from some ore. It is more probable that the bronze was prepared by smelting some ore containing tin as well as copper. For tin is but very rarely found native, and it is not very easy to separate it from any of its ores, of which *tin-stone* or *cassiterite* is the best known.

Exercise.—Endeavour to obtain tin from tin-stone (cassiterite), proceeding as in Experiment 1. If the ore is finely ground, and you maintain a high temperature for some time, you should succeed. Can you see one or two shining globules as you are heating? After you have ground the product in a mortar, you should find a number of little white shining metallic plates. Tin, therefore, is found to be a white fusible malleable metal.

Tin itself was certainly known in Egypt at a later date, for Flinders Petrie has found a small ring, dating from about 1450 B.C., which examination has shown to be of tin.

Later, at the time of the Roman Empire, tin ore was largely obtained from Western Europe. Diodorus Siculus, who wrote in Greek at the beginning of the Christian era, says:—"Tin comes from many parts of Iberia (Spain), not found upon the surface, as some in their histories have repeatedly stated, but dug up like silver and gold. There are many tin-mines beyond Lusitania throughout the islands which lie off Iberia in the ocean, which are called Kassiterides (Scilly Islands). And the ore is carried in large quantities from the island of Britain to the opposite land of Gaul, and is brought by the traders through the Keltic midlands upon horses to Marseilles and the city of Narbo." In connection with this statement it is interesting to note that up to the present day tin ore has been worked in Cornwall.

Subsequent to the extensive use of gold and copper and

bronze in Ancient Egypt, other metals gradually became known.

SILVER.—Much of the gold known to be very ancient contains silver. Gold which contains much silver is pale amber-coloured, and was often called "**electrum.**"

But silver itself was used at a fairly early date for ornaments. The diadem of a king who lived about 2700 B.C. was made of silver and gold. The beautiful jewellery of four princesses buried at Dahshur, which may now be seen in the museum at Gizeh, contains silver, besides gold and precious gems. Later, when there was more frequent intercourse with the neighbouring nations of Asia, silver became more common.

Silver occurs native, but only to a small extent. One of the ores which contains silver is known as *silver glance*.

Exercises.—1. Try whether you can separate any silver from silver glance, proceeding as in Experiment 1. What is the character of the metal? Try its hardness by pressing it in the mortar. Try its brittleness by cutting it with a knife.

2. Try to extract silver from silver residues.

LEAD came into use about the same time as silver. It was often mixed with copper and tin to make an easily fusible bronze for statuary.

The commonest ore of lead is *galena*, a heavy mineral from which the metal is readily smelted.

Exercise.—Smelt the metal from a little galena, proceeding as in Experiment 1. What is the character of the metal? It is fusible and very malleable. Try whether it will mark paper.

IRON.—Contrary to what might be supposed, iron was the last of the common metals to be brought into general use by the Egyptians. It is uncertain when it was first known in Egypt, but it appears that as late as 700 B.C. it was still uncommon enough to be acceptable as a present. Rusted wedges of iron have been found under the obelisks of Rameses II., who lived about 1300 B.C., at Tanis, near Palestine; but it is clearly not safe to conclude that they were put there in his reign, and that therefore iron was known at that time.

It may be added that it seems certain that the Assyrians and Chaldæans made use of iron sooner than the Egyptians.

One of the chief ores of iron is *red hæmatite*, an oxide of iron. In England this occurs abundantly in Cumberland, and at Furness in Lancashire.

Expt. 2. **An Attempt to extract Iron from Red Hæmatite by means of the Blow-pipe.**—Mix the hæmatite with fusion mixture, and heat it on charcoal until there is no further change. A black mass remains. Take it off with a knife, transfer it to a watch-glass, and gently crush it. It is easily broken up, and evidently is not metallic iron. Iron has a high melting-point and even if it had been set free in the metallic state might not have fused together. How can we detect a few very small specks of iron among unchanged iron oxide and specks of charcoal? The charcoal may be washed away with water. Put underneath the watch-glass a slip of white paper, and bring a magnet under this. The small specks of iron, if any, may now be detected by seeing how the magnet drags them about. Magnetise a knife-blade with a magnet, and try whether the specks are attracted and picked up by the knife-blade. If so, transfer them to a watch-glass. Add to them two drops of strong hydrochloric acid: if they effervesce, you may conclude that they were small particles of metallic iron.

As has been already stated, iron undoubtedly came into use in Egypt subsequently to the other metals which have been mentioned. A reason for this is given by the last experiment—iron is not at all readily smelted from its ores. It may, in fact, be stated that any nation which can smelt iron from its ores is relatively far advanced in metallurgical practice. At the present day far larger quantities of iron than of any other metal are produced. The ore is mixed with coal in place of charcoal, and with limestone instead of the fusion mixture we have used, and heated in huge towers, called blast furnaces, often 90 feet high. The crude iron melts, and, falling down the tower, is run off in a liquid state from a hole in the bottom into moulds. The bars so obtained are known as "pig-iron." In the huge blast furnaces, as on charcoal, the separation of iron from hæmatite is a *reduction* of an oxide to metal.

The Israelites, after their sojourn in Egypt, were, of course, acquainted with the metals used there. We read in the Old

Testament of their golden calf, of the "brazen"[1] serpent, and of the use of silver, gold, and brass in the Tabernacle. It is not till later that we read of iron. Iron is not mentioned in Exodus.

It is interesting to look up all references to metals in the earlier books of the Old Testament by means of Cruden's Concordance.

The exploration of the ruins of **Lachish** by Messrs. Petrie and Bliss since 1888 has proved particularly interesting. It was a city of the Amorites on the south-west border of Palestine. Joshua, after the victory of Bethhoron, besieged and took Lachish, and it became a stronghold of the Israelites; it subsequently saw many other changes. The great mound of Tel-el-Hesy of the present day contains the *débris* of city upon city. In the lowest layer, containing the relics of the earliest city (about 1500 B.C.), have been found rough weapons of copper. Above this have been found bronze tools, and a few pieces of silver and lead. At a higher level the chief metal was iron. This district was finally deserted about 400 B.C.

In Greece the order of the discovery of the metals followed the same course. The oldest metallic articles which are known date from 2000 B.C.: they include two small gold rings and a saw of pure copper. In the tombs of Mycenæ, the ancient capital of Achaia, the bodies were covered with ornaments of gold, and to a much smaller extent with others of silver; and in the ruins of the city, copper jugs and cauldrons, and great leaden jars for storing grain, and also elegant bronze razors, swords, and tools have been found. At Mycenæ iron was a precious metal of which finger-rings were made.

Homer, in the description of Achilles' shield in the 18th book of the *Iliad*, tells of Hephæstos, at his forge under Etna, working bars of gold and silver and copper and tin, heating them in the fire, and hammering them on his anvil into beautiful designs; and how, after the shield, he forged a cuirass of copper and greaves of tin, and a helmet with a glancing golden crest. In addition to the above metals, Homer also mentions iron. But he treats it as a rare and costly metal, an iron discus being given as a prize in an athletic contest, and he often speaks of it as being "worked with much toil."

[1] The word "brass" in the Old Testament never refers to that which we understand as brass, which is an alloy of zinc. It is used for copper or bronze.

MERCURY.—Long after the time of which Homer sings, the Greek philosopher, Aristotle, in the fourth century B.C., gave the first description of mercury. He called it *hydrarguros*, or "water silver." Indeed, on account of its liquid nature, many would not admit mercury to be a metal, and it was not until it was frozen solid (in 1760) that it was accepted as a true metal, which, unlike every other metal, is a liquid at the ordinary temperature.

Mercury is found in small globules along with its chief ore *Cinnabar*. This ore occurs in heavy red masses in Hungary, California, China, and Japan.

EXPT. 3. **To obtain Mercury from Cinnabar.**—Endeavour to obtain mercury by smelting cinnabar upon charcoal, proceeding as in Experiment 1. You fail. Why? Because mercury is volatile, and any of it which may be separated is certain to volatilise under the intense heat of the flame. Obviously, then, the ore must be heated in such a way that the mercurial vapour may be condensed.

Take a glass tube about 8 inches long with a bulb at one end (Fig. 28). This may be made from the ordinary soft glass

FIG. 28.

tubing. Mix a little finely-powdered ore as before with four times as much dry fusion mixture and a little powdered charcoal, and introduce the mixture by a paper channel into the bulb. Clean the tube with a roll of filter-paper. Warm the bulb very gently in the flame: probably some water will be driven off and condensed on the tube; absorb this, so long as it is given off, with a roll of filter-paper. Then heat the bulb rather more strongly. Look for a number of little globules of mercury deposited on the cool glass tube. Examine with a magnifying-glass. Hold the tube up to the light. The deposit looks black and opaque. Hold it in the light, but on a dark background. The deposit looks bright, reflecting the light. In this way

particles of mercury may be distinguished from other substances, for it reflects, but does not transmit, light.

Brass.—Aristotle also described a method for producing the alloy we call brass, which is a mixture of copper and zinc, by heating copper with the mineral *calamine* which contains **zinc.** Metallic zinc itself was not, however, known for another 2000 years (about 1600 A.D.).

The discovery of seven metals has been spoken of—gold, copper, tin, silver, lead, iron, and mercury. No others were generally known at the opening of the Christian era. It is doubtful whether any other metal was discovered till the end of the sixteenth century A.D. This fact may seem extraordinary, but most of the metals discovered during the last 300 years require new methods in separating them from their ores. Some other metals were not discovered at an earlier date, because the ores in which they are hidden are rare.

ORES WHICH CLOSELY RESEMBLE METALS

It is interesting to observe that the Greek word μέταλλον from which was derived the Latin *metallum* and our word *metal*, meant a "mine" or "quarry." The Latin word metallum was applied not only to what we now call metals, but to anything found in mines, *i.e.* to any minerals. There is little doubt that the mediæval belief in the transmutations of the metals was supported by this wide application of the word "metal" (*cf.* Vol. I., Chapter I. p. 2). Indeed, various ores do resemble metals very closely. We may take three examples to illustrate this.

Iron Pyrites is a widely-distributed ore of which large masses are found in Anglesea. Select a piece in which the shining blocks of ore are closely packed together and enclosed in a matrix of rock. Observe the weight of the ore, the yellow colour (a shade greyer than gold), the opacity, its crystalline form, and, above all, its bright metallic lustre.

Copper Pyrites is often closely associated with the above ore. Quantities of it are found in Cornwall. It has the colour, the opacity, and the brightly reflecting surface of brass, and it has a density which is suggestive of a metal.

Galena is a third ore of great importance, which might as readily and excusably be mistaken for a metal by a modern school-boy as by a philosopher of by-gone times. It has a leaden-grey colour, metallic lustre, and great density. Veins of galena run through the coal measures of Northumberland, Cumberland, Derbyshire, and Yorkshire.

We have previously observed the effect of heating various metals in the air (*see* Vol. I. p. 84, and Vol. II. p. 14). Let us also heat these minerals.

EXPT. 4. **To observe the Effect of roasting Iron Pyrites.** — Take a piece of hard glass tubing about 20 cm. long, and bend it slightly at a short distance from one end. Take up a few particles of the powdered ore on the point of a knife, and deposit them in the elbow of the tube. Heat the elbow, first in the luminous Bunsen flame, then in the non-luminous flame, and lastly with a blow-pipe, holding the tube as upright as possible to promote a draught. When strongly heated, observe the odour of the gases passing out of the tube. After the ore has been strongly roasted, observe the colour and appearance of the residue, and also of any sublimate which is deposited further up the tube.

Find also the effect of roasting (*a*) copper pyrites, (*b*) galena, in an elbow tube.

From each ore a gas is evolved which you recognise by its pungent odour as sulphur dioxide. A sublimate of sulphur is also obtained from the two named pyrites. The residue from iron pyrites is a dull red powder. That from copper pyrites may be purplish in colour. That from galena is white. In the case of galena you may also observe a sublimate of yellow litharge. In each case, then, a calx or earth is left, while the fumes of burning sulphur are evolved.

Two very different statements of what has occurred might conceivably be made. We might say, as some of the alchemists did, that these ores are *metals*, and that the sulphur has been separated from each of them, and that therefore sulphur is a constituent of each "metal," and therefore perhaps of all metals. And so we might conclude that by adding to or taking away from the amount of sulphur in one of these, we might change one metal into another; for example, when galena is heated on charcoal and lead obtained, we might take

the view that the galena by losing sulphur is changed into lead.

On the other hand, we might take the view of modern chemists that these ores are not metals, but substances containing sulphur along with other material, including perhaps a true metal; according to this view, when lead is obtained by heating galena on charcoal, the galena is not changed into lead, but the lead which is all the time present in the galena is separated from the sulphur which the galena also contains.

The great similarity between some of the sulphur ores and the metals proper doubtless had considerable influence in establishing the belief in the possibility of transmuting the metals, which dominated the minds of mediæval chemists. This belief led to much desultory chemical research in the pursuit of the noble metals, which, although it was of much value in leading to a knowledge of laboratory operations, retarded the true development of the science, and led chemistry to be associated with jugglery and chicanery. The alchemists even supposed that the several metals had their birth under the influence of particular heavenly bodies, and commonly denoted them by the signs used for the planets or stars. For instance, gold received the sign of the Sun, and was spoken of as Sol; silver was connected with the Moon, and iron with Mars. This astrological notion has come down to the present day in the name of Mercury for quicksilver.

The Definition of a Metal.—What characters should a metal possess? So far as we have gone, it would appear that it must possess a special lustre, and it must be hard, dense, opaque, malleable, and fusible.

Whether this definition is sufficient, or whether it may need modification, further discoveries will decide.

Problem

Find what metal, if any, can be extracted from the substances, A, B, C, D.[1]

[1] Pb_3O_4, Cu_2O, Fe_2O_3, HgO

CHAPTER VI

THE ACTION OF ACIDS ON METALS.—THE DISCOVERY OF "INFLAMMABLE AIR"

EARLY in the seventeenth century it was reported that an inflammable air had been obtained by the action of sulphuric acid upon iron. The statement, however, attracted little or no attention, and until the time of the *Hon. Henry Cavendish*, in the next century, nothing further was known of any "Inflammable Air." In the year 1766 this chemist published a paper entitled "Experiments upon Inflammable Air,"[1] in which he described the effect of treating different metals with the chief acids.

Let us set ourselves the same problem, and proceed to examine systematically the action of hydrochloric acid, sulphuric acid, and nitric acid upon the following common metals :—zinc (granulated, or in strips of the sheet metal), iron (thin pianoforte wire or filings), tin (foil or granulated), lead (sheet or granulated), copper, (turnings), aluminium (foil), and magnesium (ribbon or wire).

THE DISCOVERY OF INFLAMMABLE AIR

[A.] ACTION OF HYDROCHLORIC ACID UPON METALS

EXPT. 1. **To find the Effect of Hydrochloric Acid upon Zinc.**—Place two or three pieces of zinc in a test-tube. Add some strong hydrochloric acid, diluted with twice its volume of water. An immediate and rapid evolution of some air or gas takes place.

[1] Phil. Trans. of the Royal Society, 1766, p. 144.

Let us try whether the gas from zinc and hydrochloric acid is one of those we already know.

EXPT. 2. **Does the Gas from Zinc and Hydrochloric Acid support Combustion?**—Place some zinc in a test-tube. Add some weak hydrochloric acid. Introduce a lighted splint into the mouth of the test-tube. There is a slight explosion, and a flash travels down the tube. Repeat. The gas burns with a blue flame.

PROBLEMS

Find the effect of hydrochloric acid on magnesium, tin, iron, copper, lead, aluminium. Tabulate your results. Try first with strong acid, diluted with twice its volume of water. If there is no action, try with undiluted acid, and heat cautiously. State in each case whether any gas is evolved, and if so find whether it is an Inflammable Air.

The gas must be different from any gas we have had before, for a splint does not burn in it, as it does in atmospheric air or

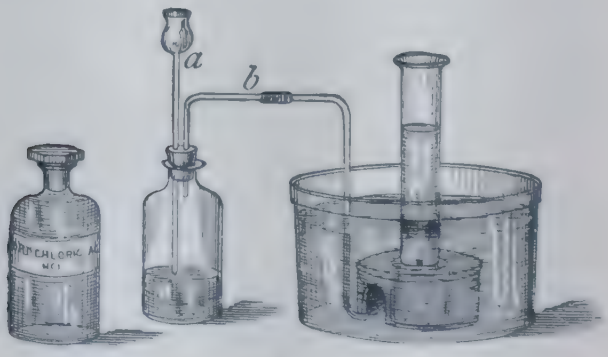

FIG. 29.

in oxygen; on the other hand the gas itself burns, being thus unlike nitrogen or carbonic acid. It was this property which led Cavendish to call it **Inflammable Air.** Find whether the gas can be collected over water (cf. Experiment 2, p. 21). Yes.

EXPT. 3. **Preparation of Inflammable Air.** Place about 20 grams of zinc in a flask provided with a delivery tube and a thistle funnel (Fig. 29). Add water so as to cover the zinc and the end of the thistle tube. Add gradually 25 c.c. hydrochloric acid previously mixed with an equal volume of water. It has been shown that at first the gas is explosive.

At this stage the gas in the delivery tube is like a train of gunpowder, the bottle like a bombshell. A light applied to the end of the tube would cause disastrous results. Collect the gas, a test-tube full at a time, over water. Try each test-tube full of it with a light. At first the gas explodes. The later portions catch fire gently, and the flame travels slowly down the tube. When it burns quietly collect the gas in larger quantities. Fill four jars. Half fill a gas jar with water, and invert it in the trough. Fill it up with the gas. It now contains air and the gas in equal parts. Set it aside.

Finally, add water down the thistle tube to dilute the acid and stop the action. Pour the liquid away. Keep the zinc that is left, after washing it with more water, for future use.

Properties of Inflammable Air

EXPT. 4. **Inflammable Air cannot burn without Common Air.**—Light a taper. Hold a jar of the gas mouth downwards, and quickly plunge in a burning splint. The gas burns quietly at the mouth of the jar, and the pale flame then travels upwards. The splint which was pushed into the centre of the jar is extinguished. As Cavendish remarked, "It appears that this air, like other inflammable substances, cannot burn without the assistance of common air."

This burning of Inflammable Air with a pale flame when only partially exposed to the atmosphere serves as a *test* to distinguish it from other gases. Is it not likely, then, that the burning will be much more violent when the Inflammable Air is *mixed* with common air?

EXPT. 5. **To burn Inflammable Air mixed with Common Air.**—Apply a lighted taper to the gas jar set aside, containing the mixture. There is a sharp explosion, or as Cavendish expressed it, the mixture of the two airs "goes off with a bounce." We now can see some reason for the precautions taken in Experiment 3.

EXPT. 6. **Is Inflammable Air lighter than Common Air?**—(*a*) Place a jar containing the gas mouth upwards on the table. Remove the lid. After one minute insert a burning splint. The content does not catch fire. It must have escaped. So it is lighter than air.

(*b*) Hold a small strong jar A mouth downwards in the air. Try to pour the gas *upwards* into it from a jar B Fig. 30). Apply a light to A. There is an explosion.

On account of the lightness of the gas it has been used for balloons. Soap bubbles blown with it will rapidly ascend through the air. Then will it not be possible to collect Inflammable Air by *upward displacement?* Suggest an apparatus. Try it.

An experiment similar to Experiment 6, p. 6, may also be tried.

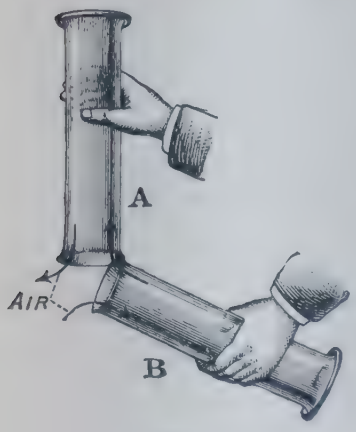

FIG. 30.

It can be shown that it is the *same* gas—"Inflammable Air"—which is obtained by the action of

(*a*) Weak hydrochloric acid on iron, magnesium, aluminium, zinc.

(*b*) Strong hydrochloric acid on tin.

Hydrochloric acid has little or no action on copper and lead.

[B.] ACTION OF WEAK SULPHURIC ACID UPON METALS.

EXPT. 7. **Do Weak Sulphuric Acid and Zinc evolve Inflammable Air?**—Place two or three pieces of zinc in a test-tube. Add some weak sulphuric acid (1 part acid, 8 parts water; see Vol. I., p. 154, Experiment 3). A gas is rapidly evolved. Apply a light. It burns. It is Inflammable Air.

PROBLEM

Find the action, if any, of cold weak sulphuric acid (1 : 8) upon aluminium, magnesium, iron, tin, copper, lead.

Inflammable Air is commonly prepared by the action of weak sulphuric acid upon zinc in a way similar to that described in Experiment 3.

Precautions in the Preparation of Inflammable Air.—(1) The gas is often made in a Woulff's bottle of thick glass.

Though this is not readily broken by heavy pieces of zinc falling into it, thick glass is specially apt to crack on heating. It is important, therefore, to keep the contents *cool*.

(2) After mixing 8 parts of water with 1 part of sulphuric acid, set the mixture aside to cool; otherwise it might crack the Woulff's bottle. It is better still to keep a stock of diluted acid ready.

(3) Make sure that all air is expelled before the gas is collected (*see* p. 66, Experiment 3).

(4) Should the contents of the bottle become hot, the current of gas will probably slacken. It is then best to pour off the liquid, wash the zinc, and start afresh.

(5) Take care not to breathe the gas made from ordinary impure zinc.

(6) The addition of a few drops of a solution of copper sulphate often quickens the current of gas.[1]

EXPT. 8. **To find what becomes of the Zinc in the Preparation of Inflammable Air.**—(1) Zinc placed in excess of dilute sulphuric acid entirely disappears. Can it be recovered? Fill a beaker one-third full with dilute sulphuric acid. Put in several pieces of zinc. A vigorous effervescence goes on and gradually subsides. Why does the effervescence cease? Is the sulphuric acid all spent? Or the zinc all used up? Add more zinc. Effervescence begins again. This shows there is still some acid in the liquid. Go on adding zinc from time to time until there is no more action, and the sulphuric acid has spent itself. The liquid now contains undissolved zinc, and floating black specks, which are impurities left by the dissolved zinc. Filter. Evaporate over wire gauze, and try to crystallise.

Not zinc, but long white crystals are obtained. Examine them. Their crystalline form, their "salty" rather bitter taste, and the readiness with which they dissolve in water (the solution is shown by litmus to be slightly acid) may lead us to class the substance with nitre, with sal-ammoniac and others, which we have called *salts* (cf. Vol. I., p. 167). We may in fact regard it as a salt, prepared by a new method, and we may suitably call it **zinc sulphate.**

Try the effect of heating a small quantity of the substance in

[1] *See* Appendix for a description of Kipp's Hydrogen Apparatus.

a dry test-tube. It contains water of crystallisation, and a white amorphous residue is left. Determine the percentage of water of crystallisation, proceeding as in Vol. I., p. 141, Experiment 12.

It is found to contain 44 per cent.

(2) Repeat with hydrochloric acid and zinc. You will have to evaporate to a very small volume, and even then may not succeed in obtaining any crystals. Evaporate some of the clear liquid on a watch-glass upon a sand-bath. The white non crystalline solid at last obtained is the salt **zinc chloride**. This substance, when exposed to the air, rapidly becomes liquid through attracting the moisture in the air, and dissolving in it. Such a substance is said to be *deliquescent*.

Problems

Prepare the following salts and describe their appearance:

(1) Sulphates of magnesium, iron, nickel.
(2) Chlorides of magnesium, tin, iron.

What is the Source of Inflammable Air?—Does the Inflammable Air come from the metal? We have obtained it by using several metals; then, if it comes from the metals, they are not elements. Or does it come from the acid? If so, more than one acid contains it as a common constituent? Or does the Inflammable Air come from both the metal and the acid? Does it perhaps come from the water? And is it an element or a compound?

We cannot as yet answer these questions. We must wait for further evidence, and must in the meantime endeavour to avoid stating that the gas is obtained from the metal.

Measurement of the Volume and Weight of Inflammable Air obtainable by the Action of Acids on Metals.

Expt. 9. **To find the Volume of Inflammable Air obtainable by means of 1 Gram of Magnesium.**—Fit up the apparatus shown in Fig. 31, consisting of funnel, rubber tube, burette, and wide glass dish or beaker. The rubber tube *b* must fit the burette tightly. Take 5 cm. of magnesium ribbon. Its weight will be only a few centigrams.

and a single milligram will make a large percentage difference in the result. Weigh, if possible, exactly to a milligram. Disconnect the rubber tube from the burette. Gently bend the magnesium. Push it one-third of the way up the burette to *a*. Suck water up the burette till it nearly reaches the top of the graduations. Close the tap. Re-attach the rubber tube. Support the whole by the ring of a retort-stand. Half fill the funnel with dilute hydrochloric acid. Read the level of the water. Gently open the tap. Acid runs through. As soon as bubbles start from the magnesium turn off the tap. The magnesium is quickly and almost entirely dissolved. The burette must now be slightly tilted and a few more drops of acid skilfully run down the side to dissolve the rest of the metal. Again read the position of the liquid in the burette. Subtract one reading from the other.

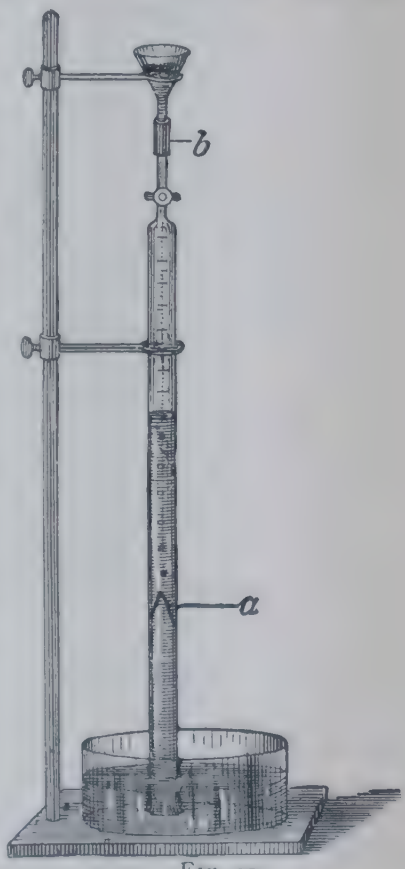

FIG. 31.

Thus, in an experiment,

·025 gram magnesium gave with acid 23 c.c. Inflammable Air.

Hence, 1 gram magnesium would give with acid $\frac{23}{\cdot 025}$ or **920** c.c. Inflammable Air.

Do *you* find that equal weights of magnesium give with acid the same volume of gas?

Exercise.—Weigh accurately about 0·1 gram of zinc foil in a thin strip. Find the volume of Inflammable Air obtained by the action of hydrochloric acid upon it.

EXPT. 10. **To find the Volume of Inflammable Air obtainable by means of 1 Gram of Zinc.**—Fit up the apparatus shown in Fig. 32, consisting of a 4-oz. flask, Winchester pint bottle, beaker, rubber corks, glass tubes, rubber junctions, and clip. The rubber corks *must* fit the flask and the bottle

tightly. Nearly fill the bottle with water. Open the clip *c*. Suck at the end of the tube *d* so as to fill it with water, and then close the clip *c*. The tube will remain full of water. Place an empty beaker under *d*. Fill the flask *b* half full of dilute sulphuric acid (1 of acid to 3 of water).

Weigh out accurately about 1 gram of zinc foil. Remove the flask *b* attached to the cork, introduce the zinc, and *at once* replace the cork, and open the clip *c*. The gas evolved as the metal is dissolved enters the Winchester bottle, and forces an

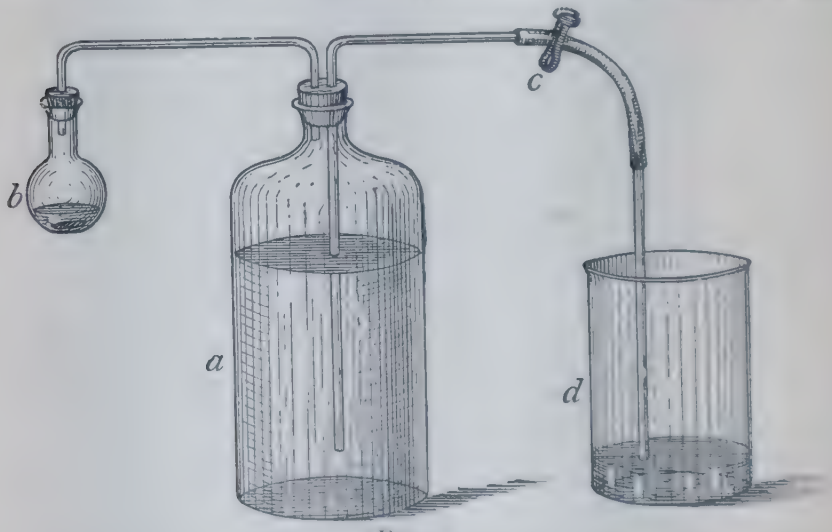

FIG. 32.

equal volume of water through the tube into the beaker. If the solution of the metal slackens or stops, the flask *b* must be heated over wire gauze.

When the metal is completely dissolved the flask *b* and its contents are found to be hot, and must be cooled. Why? This can be readily done by placing the flask in a tumbler of cold water, and gently shaking the flask so as to cool its sides and the gas within it. Next, raise the beaker *d* until the level of the water in it is the same as that of the water in the bottle. Then close the clip *c*. Now measure the volume of the water in the beaker by pouring it into a measuring jar. This is also the volume of gas evolved on the solution of the metal.

Thus, ·95 gm. zinc gave with acid 350 c.c. Inflammable Air.

Hence, 1 ,, ,, ,, $\frac{350}{\cdot 95}$ c.c. ,,

$= 368$ c.c. ,,

Problems

Find the volume of Inflammable Air obtainable from 1 gram of

(1) Zinc with dilute hydrochloric acid (1 of acid to 2 of water).

(2) Iron with (*a*) dilute sulphuric acid (1 of acid to 3 of water), (*b*) dilute hydrochloric acid. It may be necessary to heat the flask *b*.

(3) Magnesium with (*a*) dilute sulphuric acid, (*b*) dilute hydrochloric acid. Wrap the metal in paper. Take about 0·25 gram metal.

(4) Tin with strong hydrochloric acid. It will be necessary to heat the flask *b*.

(5) Aluminium with dilute hydrochloric acid. Take about 0·25 gram metal.

Can you sum up your results in one or two general statements?

Results.—Compare your results with the following *mean* results obtained by a class of boys for the volumes of Inflammable Air obtainable from 1 gram of metal and hydrochloric or sulphuric acid.

	Hydrochloric Acid.	Sulphuric Acid.
Tin	224 c.c.	—
Zinc	378 ,,	376 c.c.
Iron	453 ,,	454 ,,
Magnesium...	1008 ,,	1002 ,,
Aluminium	1340 ,,	—

What do these results show? Do not the numbers suggest the conclusion, to which Cavendish came, that

(*a*) the *same* volume of Inflammable Air was obtained whether sulphuric or hydrochloric acid was used,

(*b*) the same weight of different metals gave *different* volumes of Inflammable Air.

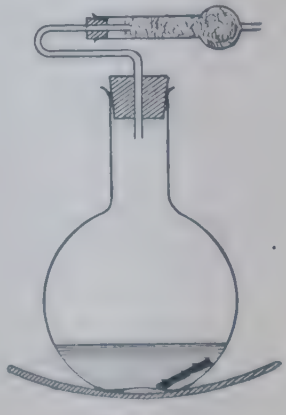

Fig. 33.

*Expt. 11. **To find the Weight of Zinc by means of which 1 Gram of Inflammable Air can be obtained.**—Fit up the apparatus shown in Fig. 33, consisting of a 4-oz. flask, "drying tube," cork,

and bent glass tube. The flask must be light, and of thin glass. Pack the tube with dry "calcium chloride." Fill one-quarter of the flask with dilute sulphuric acid. Add one or two drops of a solution of copper sulphate. Weigh the whole on a good balance. Weigh out accurately 0·5 to 1 gram of pure zinc foil. Remove the cork, introduce the zinc, and at once replace the cork. If the evolution of gas becomes rapid, the flask must be cooled by placing it in cold water. Calcium chloride attracts water, and its purpose is to prevent the escape of water with the gas. When the zinc is completely dissolved, dry the outside of the flask with a cloth and weigh again. Thus :—

	Weight of flask and acid	= 98·689 gm.
	Weight of zinc	= 0·971 ,,
Then,	Weight of flask, acid, and zinc	= 99·660 ,,
	Final weight of flask, &c.	= 99·628 ,,
Hence,	Weight of Inflammable Air evolved ...	= ·032 ,,

Thus,

·032 gram Inflammable Air was obtained by means of 0·971 gram zinc.

Hence,

1 gram Inflammable Air was obtained by means of 30·3 grams zinc.

It will probably be difficult to obtain an accurate result on account of the very small weight of the Inflammable Air. Still, it was in almost exactly this way that Cavendish tried to answer the same question.

The Weight of 1 Litre of Inflammable Air can now be calculated. For we have found that—

1 gm. Zinc and acid liberate 376 c.c. Inflammable Air.

∴ 30·3 ,, ,, ,, ,, $376 \times 30{\cdot}3$ c.c. ,,

$= 11{,}392{\cdot}8$ c.c. ,,

$= 11{\cdot}4$ litres ,,

But 30·3 gram Zinc liberate 1 gm. ,,

∴ 11·4 litres of Inflammable Air weigh 1 gm.

∴ 1 litre ,, weighs $\frac{1}{11{\cdot}4}$ gm.

$= 0{\cdot}09$ gm.

The gas is, therefore, extraordinarily light. The density of atmospheric air is about 14 times as great.

Problems.

1. Collect **Inflammable Air by upward displacement.**

2. Try to burn Inflammable Air at the end of a delivery tube; observe all necessary precautions. Obtain some idea of the temperature of the flame by observing its effect on thin iron and copper wires.

CHAPTER VII[1]

THE ACTION OF ACIDS ON METALS.—THE DISCOVERY OF VOLATILE SULPHUROUS ACID

ACTION OF STRONG SULPHURIC ACID UPON METALS.—This is the next question to be considered in continuing our research upon the action of acids upon metals.

EXPT. 1. **To find the Effect of Strong Sulphuric Acid upon Zinc.**—Place a piece of zinc in a test-tube. Cover it with strong sulphuric acid. A few bubbles are given off. Warm. The bubbling increases. A gas is given off. Test it with a light. A splint is extinguished in it, and the gas does not catch fire. It is not Inflammable Air. It has a strange pungent sulphurous smell. A new gas is being formed. (Allow the acid to cool before pouring it into cold water and throwing away.)

PROBLEM

Find the effect of hot strong sulphuric acid upon copper, iron, tin, mercury, lead.

Cavendish observed in 1766 that on heating zinc, iron, or tin with strong sulphuric acid, there was given off what he called **Volatile Sulphurous Acid.**[2] He did not collect it, nor examine it. This was first done by Priestley in 1775. He collected the gas over mercury (we shall see presently why he did not collect it over water). Since the gas was obtained

[1] Some teachers will prefer to proceed at this stage with the study of the Composition of Water (Chapter IX.), before pursuing further the lines of Cavendish's research on the Action of Acids upon Metals (Chapters VII. and VIII.), which does not lead to a study of substances so important as Inflammable Air.

[2] It is now known as sulphur dioxide.

by the action of strong "vitriolic acid" on metals, Priestley called it *vitriolic acid air*.

It is best prepared by means of copper and strong sulphuric acid.

EXPT. 2. **Preparation of Volatile Sulphurous Acid.**— Introduce 10 grams of copper into an 8-oz. flask. Fit it with a stopper through which are passed a thistle funnel and a delivery tube bent twice at right angles (Fig. 34). Support the flask

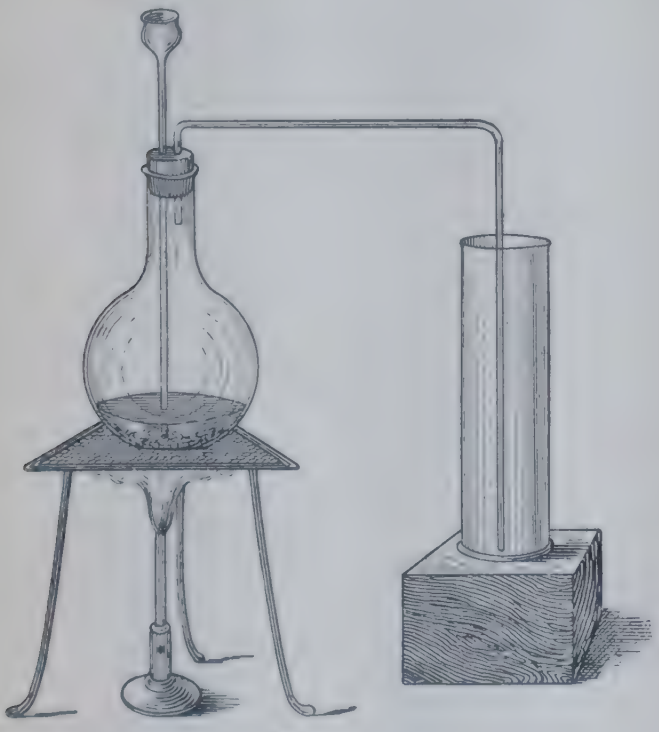

FIG. 34

upon wire gauze on a tripod stand, and place a *dry* jar so as to collect the gas by downward displacement. Add 50 c.c. of strong sulphuric acid. Heat cautiously. Plunge a burning splint from time to time into the jar. As soon as it is extinguished at the mouth of the jar, remove the jar, closing it with a stopper or glass plate, and substitute another *dry* jar. Fill three jars.

Half fill a test-tube with water. Pass the current of gas to the bottom of it. Notice that the bubbles of gas do not reach the surface. The gas must be soluble in water. (Water, at the

ordinary temperature, can absorb about 40 times its own volume of the gas.) After passing the gas into the water for five minutes, disconnect the flask and place it in a fume closet.

Properties of Volatile Sulphurous Acid.

We have observed that the gas is not inflammable, and that it does not support combustion. It is heavier than air (for we were able to collect it by downward displacement). It possesses a suffocating sulphurous odour, unlike any other gas we know. Why could we not collect it over water?

EXPT. 3. **Is it soluble in Water?**—Invert a jar filled with the gas in water, and withdraw the glass plate closing its mouth. Shake the jar so as to wet the sides. The water *ascends* in the jar. Why? The gas is certainly soluble in water. If there be any residue left in the jar after shaking, find out what it is.

EXPT. 4. **Properties of the Aqueous Solution of the Gas.**—Divide the liquid into three parts :—

(*a*) Smell it ; it has the same odour as the gas. Taste it : it is acid.

(*b*) Add blue litmus ; it is reddened.

(*c*) Add lime-water ; there is no turbidity.

EXPT. 5. **To show that the Gas bleaches many Colouring Matters.**—(*a*) Pour a dilute solution of the dye magenta into a jar containing Volatile Sulphurous Acid. Cover the mouth with the hand, and shake well. The colour is nearly or entirely destroyed.

(*b*) Place some red or blue flowers in a jar of the gas. The colours are soon destroyed.

The gas is used for bleaching straw, silk, and woollen goods.

We must for the present leave this bleaching action unexplained.

Volatile Sulphurous Acid is found also to be a disinfectant. Rooms, bedding, and clothing which contain disease germs are therefore often exposed to the action of the gas.

EXPT. 6. **To find what becomes of the Copper in the Preparation of Volatile Sulphurous Acid.**—When the mixture in the flask (Experiment 2) is cold, pour the liquid slowly and as completely as possible into water. Examine the cake which remains in the flask. Add a little water. Some small blue crystals can probably be seen. Warm,

so as to dissolve the crystals, and filter the liquid into a crystallising dish, and allow it to cool. Beautiful blue crystals are deposited, which may be dried upon filter-paper. What are they?

(2) You may have noticed accidentally that if a knife is moistened with the blue solution, the steel is reddened. Confirm thus: dissolve some of the blue crystals in a little water, and dip a clean steel knife-blade into the blue solution. A copper-coloured deposit is at once formed on the steel. It is evident, therefore, that *the blue crystals contain the copper*, and we may name them **sulphate of copper.**

Their similarity to crystals of Blue Vitriol is evident, and by determining their percentage of water of crystallisation, and by other tests, they may in fact be proved to be crystals of Blue Vitriol.

PROBLEM

Find what gas is given off by the action of strong sulphuric acid upon (*a*) magnesium, (*b*) aluminium.

The Composition of Volatile Sulphurous Acid.—Priestley quite expected that when Volatile Sulphurous Acid was dissolved in water sulphuric acid would again be obtained. But, on the contrary, as he himself wrote in 1775:—

"Water impregnated with Volatile Sulphurous Acid differs remarkably from sulphuric acid.[1] Its acidity is now become trifling to what it was, and from being the most fixed and strongest, it is now become the weakest and the most volatile of all acids the smell of it being intolerably pungent, and almost the whole of it evaporating when it is exposed to the open air."

It is therefore evident that the gas does not occur in solution in ordinary sulphuric acid.

Let us call the acid which is formed when the gas is dissolved in water **sulphurous acid**, and so distinguish it from sulphur*ic* acid.

The odour of Volatile Sulphurous Acid is very similar to that produced when sulphur is burnt in oxygen or in air. Why? Perhaps the *same* gas is produced by burning sulphur as by the action of strong sulphuric acid upon metals.

EXPT. 7. **To find whether the Burning of Sulphur in**

[1] His own words here are "vitriolic acid air" and "vitriolic acid."

Oxygen also produces Volatile Sulphurous Acid.—Place some sulphur in a deflagrating spoon. Light it and plunge it into a jar of air. Test the properties of the contents of the jar, and compare them with those of Volatile Sulphurous Acid. They are *the same*.

But what is the composition of the fumes of burning sulphur? According to the Theory of Combustion advanced by Lavoisier, they are produced by the union of the oxygen of the air with sulphur. And if this is so, the gas liberated by the action of strong sulphuric acid upon copper and other metals must also be **a compound of sulphur and oxygen**, for we have just proved it to be identical with the fumes of burning sulphur. It must be **an oxide of sulphur**, and not an elementary substance.

In that case what is the *source* of the oxygen, and of the sulphur in the gas Volatile Sulphurous Acid when prepared by the action of oil of vitriol upon copper? The copper? or the oil of vitriol? Now, oil of vitriol can be prepared from green vitriol, and green vitriol can be obtained by exposing iron pyrites to the air and moisture, and we know that iron pyrites contains sulphur (see Vol. II., p. 63). It is, therefore, quite possible that oil of vitriol contains sulphur. Moreover, the copper cannot have been decomposed, for it has been proved that the copper could be recovered from the crystals of blue vitriol, which were obtained. It is, therefore, almost certain that the sulphur and oxygen in Volatile Sulphurous Acid are derived from the oil of vitriol, and that this acid contains sulphur and oxygen. And we can now understand why oil of vitriol may be called *sulphuric* acid.

Problems

1. Find whether any air disappears when some sulphur, melted and ignited in an iron spoon, is poured into a dry flask, which is then immediately stoppered.

2. Find whether the products of burning sulphur are heavier than the original sulphur. Place about 0·3 gram sulphur in the bulb of a calcium chloride tube, and attach the other end by a rubber cork to another similar tube packed with soda lime. Arrange to suck a slow current of air, as in Experiment 9, p. 19, over the sulphur and through the second tube. Weigh the two tubes separately, both before and after the combustion.

CHAPTER VIII

THE ACTION OF ACIDS ON METALS—THE DISCOVERY OF NITROUS AIR

ACTION OF NITRIC ACID UPON METALS.—It remains to try the action of nitric acid upon metals.

EXPT. 1. **To find the Effect of Cold Nitric Acid upon Copper.**—Place about 1 gram of copper in a small flask. Mix 5 c.c. strong nitric acid with 5 c.c water. Cool the mixture. Pour it into the flask. Is there any action? There is an effervescence. Copious red fumes are evolved. Test it with a lighted splint. It is extinguished, and the gas does not burn. There is a characteristic smell. Red fumes appear. Yet *another new gas* is being formed.

PROBLEM

Find the effect of nitric acid (1 to 1) upon iron, tin, lead.

You have been repeating experiments which were made more than 100 years ago. Cavendish, in 1766, tried the action of nitric acid upon zinc, iron, and tin. When giving an account of his experiments, all he wrote was :—

"All these metallic substances dissolve readily in the nitric acid and generate air, but the air is not at all inflammable."

Priestley was the first to collect the gas and examine its properties. This is what he wrote :—[1]

"On the 4th of June, 1772, I first found this remarkable

[1] *Experiments and Observations on Different Kinds of Air*, Dr J. Priestley, 1790 edition, vol. 1, p. 328.

species of air. . . . When I first began to speak and write of it to my friends, I happened to distinguish it by the name of **'Nitrous Air'** [1] because I had procured it by means of spirit of nitre."

Priestley wrote that he also obtained Nitrous Air by the action of nitric acid upon iron, tin, and lead, but none in the case of zinc. Further, he adds :—

"I have generally found it most convenient to get Nitrous Air from copper, on account of the pretty equable solution of the metal in the nitrous acid. If iron is made use of, the process is much more difficult."

The method of preparation by means of copper is still commonly used.

EXPT. 2. **Preparation of Nitrous Air.**—Fit up the apparatus in Fig. 35, with a glass delivery tube. Place 30

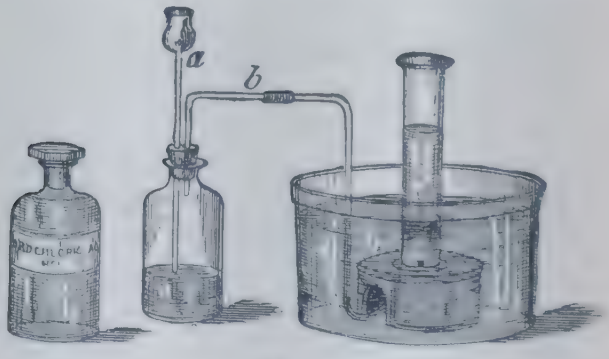

FIG. 35

grams of copper turnings in the flask. Mix 50 c.c. strong nitric acid with 50 c.c. water. Cool. Pour the acid down the funnel so as to cover the end of the tube. Red fumes are produced. At first little gas comes over. Then a colourless gas is collected. Neglect the first two jars. Collect three more. If the action slackens, add 1 or 2 c.c. strong nitric acid.

PROBLEMS

1. What became of the copper in the last experiment? Try to obtain crystals of copper nitrate. (Cf. p. 69, Experiment 8.)

2. Try to prepare crystals of the nitrates of (a) zinc, (b) iron, (c) lead.

[1] It is now known as Nitric Oxide.

Properties of Nitrous Air.

Nitrous Air is a colourless gas, insoluble in water, and we found it did not support the combustion of a taper. Let us try this again.

EXPT. 3. **Does Nitrous Air not support Ordinary Combustion?**—Plunge a lighted splint into a jar of the gas. It is extinguished.

PROBLEM

Find whether sulphur and phosphorus kindled in a deflagrating spoon will burn in nitrous air.

Observe that red fumes always appear about the mouth of the jar when it is exposed to the air, having a characteristic disagreeable smell.

What is the cause of these red fumes? Is it not probable that the red fumes are produced by the Nitrous Air mixing with common air? Let us test this.

EXPT. 4. **Are the Red Fumes formed by Nitrous Air mixing with Common Air?**— Fill a jar one-third full with water; invert it over water so that it is two-thirds full of air. Bubble up into it one-third of a jar of the Nitrous Air. Red fumes at once appear. Shake. The water gradually rises, and at the same time the red fumes disappear. Finally the two parts of common air and one part of Nitrous Air yield rather less than two parts of gas.

What has become of the rest? It must have dissolved in the water. Push a strip of blue litmus paper up into the water in the jar. It is reddened. Then the red fumes have dissolved in the water, forming an *acid* solution. Here is what Priestley wrote of the experiment:—

"I hardly know any experiment that is more adapted to amaze and surprise than this, which exhibits a quantity of air which as it were devours a quantity of another kind of air half as large as itself, and yet is so far from gaining any addition to its bulk that it is considerably diminished by it."

Have we completely determined the cause of the red fumes? No. Common air contains both oxygen and nitrogen, so we must find whether the red fumes are due to the action of the nitrogen or of the oxygen, or of both, upon the Nitrous Air.

EXPT. 5. **Does Nitrous Air form the Red Fumes with Nitrogen?**—Obtain some nitrogen as in Experiment 8. p. 18. Open the flask under water, and shake. Bubble the nitrogen up into a jar (standing over water) containing a little Nitrous Air. *No* red fumes result, and the volume of gas is increased.

EXPT. 6. **Does Nitrous Air form Red Fumes with Oxygen?**—Fill two-thirds of a jar with oxygen. Add to it one-third of a jar of Nitrous Air, as in Experiment 4. Priestley first tried this experiment in 1775. He found that the volume was diminished and "the redness was really deeper and the diminution something greater than common air would have admitted."

Add another third of a jar of Nitrous Air. There is a further diminution in volume.

Then, the red fumes produced when Nitrous Air is added to common air, are due to the Nitrous Air forming a red coloured gas with the *oxygen* of the air, which red gas is readily dissolved in water, forming an acid solution.

Nitrous Air may be used as a Test for Oxygen.—Since Nitrous Air forms red fumes with oxygen and not with any other gas, we are provided with a *test* to see whether a gas is oxygen, or has oxygen mixed with it.

Further, from the amount of contraction when the gases are shaken with water, we can form some idea of the amount of oxygen present. This test was formerly largely used by chemists after its discovery by Priestley. Cavendish showed, however, that it is not possible to obtain accurate results with simple apparatus, nor without observing many points of detail.

John Dalton was accustomed to analyse air in this way. On giving, in 1802, an account of his experiments, he wrote :—[1]

"In all parts of the earth and at every season of the year the bulk of any given quantity of atmospheric air appears to be reduced nearly 21 per cent., by abstracting its oxygen. Air brought from the summit of Helvellyn in Cumberland [1,100 yards above the sea] gave no perceptible difference from the air in Manchester."

[1] *See* Alembic Reprint, No. 2, pp. 8 and 14.

What is the Composition of Nitrous Air?—Is it a compound or an element? We found that vigorously burning phosphorus would continue to burn in Nitrous Air (Experiment 3, Problem). This suggests that Nitrous Air may contain oxygen. How can we make sure? Refer back to p. 45, Experiment 11. Let us see whether a metal, magnesium, for example, can extract oxygen from Nitrous Air.

EXPT. 7. **To find whether Nitrous Air contains Oxygen.**—Place 20 gm. of copper turnings in a bottle, connected as shown in Fig. 36 with a small flask, *c*, containing some

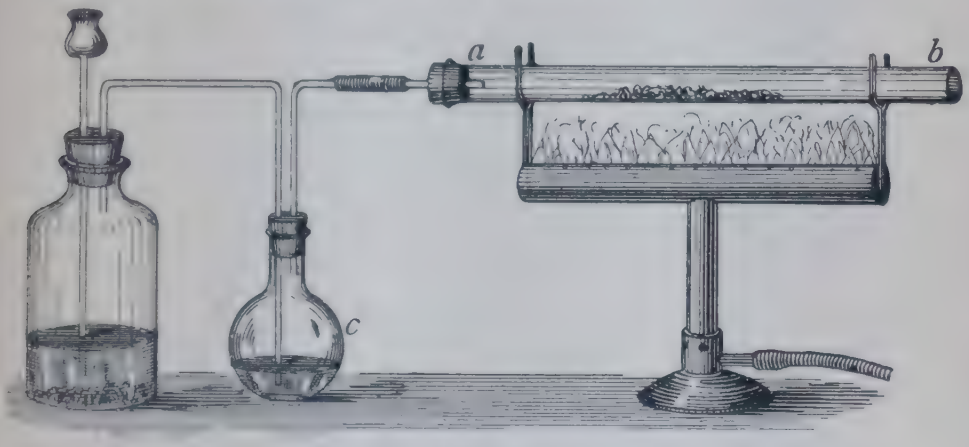

FIG. 36.

strong sulphuric acid, the tube by which the gas will enter the flask dipping just below the surface of the acid. The sulphuric acid in *c* will dry the Nitrous Air, and also indicate how fast the gas is being generated. Take a hard glass tube, *ab*, about 12 inches in length, and place in it about 100 cm. of magnesium ribbon twisted into a loose tangle. Support the tube over a Ramsay burner, and connect it with the flask *c*. Attach to the end *b* a delivery tube dipping into water in a tumbler. Cover the copper with water, and add strong nitric acid until Nitrous Air is liberated in a steady stream. When it has swept the air before it out of the apparatus, light the burner, and presently turn the coal gas on full. After a time the magnesium catches fire and burns brilliantly. Collect the gas which issues from the delivery tube, while the magnesium is burning, in test-tubes. When the action is over, examine the contents of the tube *ab*, and also the gas collected.

The tube contains a white calx, identical with the oxide of magnesium. Undoubtedly, therefore, *Nitrous Air contains Oxygen.*

The gas in the test-tubes no longer reddens on contact with the air. Apply a light. It goes out. Is it carbonic acid gas? Add lime-water. It remains clear. Then the gas is *nitrogen.*

Nitrous Air, therefore, contains both nitrogen and oxygen. But so does atmospheric air. Is it the same as common air? No, for it extinguishes a splint, and it forms red fumes with air or with oxygen. Then, in what does the difference between common air and Nitrous Air consist? This question must be left unanswered at present.

Nitrous Air is now called **Nitric Oxide**, a name which denotes the fact that it is a compound of nitrogen and oxygen. Observe that the name would be equally applicable to atmospheric air, so far as we know at present.

Problem

Find the effect of passing Nitrous Air over heated (*a*) copper, (*b*) iron.

The Action of Aqua Regia upon Metals

Every metal we have tried has dissolved in at least one or other of the acids, nitric, sulphuric, or hydrochloric. Can gold be dissolved by these acids?

Expt. 8. **Gold is soluble in Aqua Regia only.**—Half fill a test-tube with strong hydrochloric acid. Into a second test-tube pour one-third of this volume of strong pure nitric acid.

(1) Breathe on a glass rod. Take up half a sheet of gold leaf and put it in the nitric acid. Warm gently. Stir. It breaks up into little flakes but does not dissolve.

(2) Repeat with the hydrochloric acid. The gold is again not attacked.

(3) Mix the liquids. The gold dissolves readily.

This mixture of hydrochloric and nitric acids is called **Aqua Regia**, because it dissolves the royal metal gold.

Problems

1. Find what acids will dissolve (*a*) platinum, (*b*) silver.
2. Find whether the given { coin / metal / gilt leaf } is gold or tinsel.

Exercises[1]

Preparation of Gases

Try the following experiments, at first in test-tubes. Examine the properties of the gases formed. You are not expected to identify them in every case. Devise suitable apparatus for repeating the experiments on a larger scale.

1. Heat a mixture of bleaching-powder, water, and copper oxide.
2. Heat strong sulphuric acid and powdered charcoal.
3. Heat strong nitric acid and a strong solution of ferrous sulphate.
4. Heat a strong solution of caustic soda with some zinc and a few drops of copper sulphate solution.
5. Heat zinc and dilute nitric acid—

   (*a*) 40 c.c. acid + 160 c.c. water,
   (*b*) 20 c.c. acid + 180 c.c. water.

6. Heat solutions of ammonium chloride and caustic soda.
7. Heat solid ammonium chloride and strong sulphuric acid.
8. Heat ammonium nitrate crystals.
9. Heat lead nitrate crystals.
10. Heat oxalic acid crystals and strong sulphuric acid.
11. Add strong sulphuric acid to sodium formate crystals.
12. Find out what is produced by the combustion of the gas obtained in No. 11. Do you obtain any evidence of the composition of the gas?
13. Add a saturated solution of sal-ammoniac to a stick of potassium nitrite, about 3 inches long, and then heat.

[1] *To Teachers.*—These riders are important. They give an opportunity for the exercise of skilful manipulation, of accurate observation, and of sound judgment.

CHAPTER IX

THE DISCOVERY OF THE COMPOSITION OF WATER

(A) THE SYNTHESIS OF WATER

WHEN Inflammable Air is mixed with common air and fired, it explodes, or as Cavendish put it, it goes off "with a bounce." What will happen if it is mixed with oxygen instead of common air?

EXPT. 1. **To find whether Inflammable Air explodes much more violently with Oxygen than with Common Air.**—Fill a sound soda-water bottle with water; invert it in water. Bubble Inflammable Air up into it, until two-thirds of the water is displaced. Fill it up with oxygen. Close the neck with a cork. Wrap a cloth round the bottle. Remove the cork and quickly hold the neck to a flame.

This experiment was first tried by Priestley about 1774, and he remarked, "The effect far exceeded my expectations, and it has never failed to surprise every person before whom I have made the experiment. The report will be almost as loud as that of a small pistol; being, to judge by the ear, not less than 40 or 50 times as loud as with common air."

Can you suggest any reason for the explosion being so much more violent in the one case than the other? The explosion is so rapid that we can hardly say what has occurred. Let us try to burn Inflammable Air *slowly* in a closed vessel containing common air or oxygen.

EXPT. 2. **To find whether a Contraction of Volume occurs when Inflammable Air is burnt in Common Air or in Oxygen.**—Select a wide-necked flask *a* of about

500 c.c. capacity (Fig. 37). Fill a large deep trough or bucket with warm water. Place some zinc in a small (2 oz.) flask *b* and add dilute sulphuric acid, so that Inflammable Air is rapidly evolved. Close it with a cork provided with a short upright delivery tube *c*, the end of which has been slightly narrowed by holding it in a Bunsen flame. Light the Inflammable Air at the opening of the tube, observing the necessary precautions (*see* p. 69). Hold the small flask *b* in the water, with the flame just above the surface, and invert the larger flask *a* over the flame, so that its mouth is under the water. Then, releasing the flask *b*, the upward pressure of the water will cause it to take the position shown.

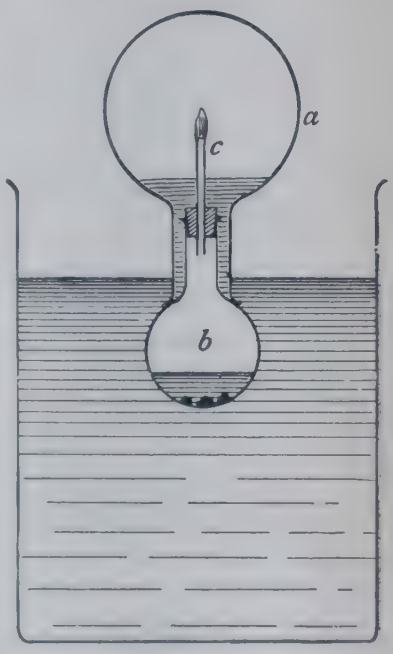

FIG. 37.

The Inflammable Air continues to burn, but the flame soon becomes fainter and smaller. As soon as it is seen to be extinguished quickly remove the flask *b* by lowering it in the water. Observe *that the water has risen in the flask a.* Then, some of the air has disappeared. Gently shake the flask *a* to throw the water against the top, in order that the flask and the residual air may acquire the temperature of the water. Close the mouth of the flask *a* by a cork under water. Remove the flask *a* and pour its contents into a measuring cylinder. Find also the total capacity of the flask *a.*

To quote an actual experiment :—

Volume of water which entered the flask	105 c.c.
Total volume of the flask	490 c.c.

Hence,

Contraction in the volume of the air in the flask = **21·4** per cent.

You also have probably found a contraction of the air. How much ?

Problem

Repeat the experiment with the difference that a longer delivery tube is used, and the flask is filled with oxygen in place of air. Find what volume of the oxygen is lost.

These experiments were made by Scheele (about 1773). He found that on using common air, the *fifth* part of it was lost. In the case of oxygen, nearly the *whole* was lost.

In each case the oxygen present seems to disappear. What has become of it? Unless our Theory of Combustion and that of the Conservation of Mass is at fault, the oxygen cannot have been destroyed. Has some unseen vapour dissolved in the water? Suppose we observe the burning of the gas in a clean dry test-tube.

Expt. 3. **To find whether a vapour is formed when Inflammable Air is burnt.**—Take a clean dry test-tube. Introduce a little Inflammable Air into it by upward displacement. Hold its mouth to a flame. It explodes. Carefully examine the sides of the tube.

Priestley did this, and found that "the inside of the glass though clean and dry before, immediately became *dewy*."

Where has the dew come from? Is it merely the moisture which we know is present in common air?[1] If so, we still have not accounted for the Inflammable Air and oxygen which disappeared in Experiment 2. It is evidently important to see whether there is after all any *loss of weight*. This task the Hon. Henry Cavendish set himself, when he heard of Priestley's experiment.

Cavendish's Experiment to find whether there is a Loss of Weight on firing a Mixture of Common and Inflammable Air.[2]—We cannot readily repeat Cavendish's experiment. He used a strong glass globe. He filled it several times with different proportions of common and Inflammable Air. The mixture was fired by electricity. He found *no loss of weight*. In all the experiments the inside of the glass globe became dewy. He also succeeded in measuring the diminution in volume produced by the explosion.

[1] This is the conclusion of Priestley and Watt (see p. 91). See *Experiments on Air*, by the Hon. Henry Cavendish, Alembic Club Reprint, No. 3, p. 12.

[2] *Loc. cit.*

Cavendish concluded that "when inflammable and common air are exploded in a proper proportion, almost all the inflammable air and near *one-fifth* (*see* Experiment 2) of the common air, lose their elasticity and are condensed into dew."

What is the nature of this dew? How can we obtain a larger amount of it? Perhaps by burning the gas under a cold surface.

EXPT. 4.[1] **To collect the Dew formed on burning a Mixture of Common and Inflammable Air.**—Fit the tubulure of a retort *a* with a cork, through which is passed a piece of glass tubing *b*, as shown in Fig. 38. Fix the retort in

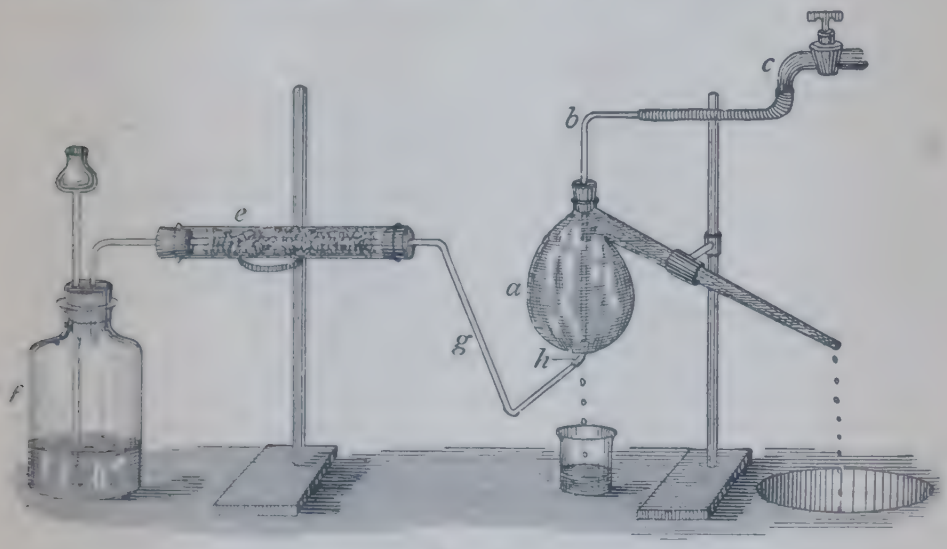

FIG. 38.

a clamp, and connect the tube *b* with a water tap *c*. On turning the tap, water flows into the retort, fills it, and drips out from the neck. Place the retort so that the water drips into a sink.

We are going to burn Inflammable Air. But we must be sure that it is free from moisture. This can be secured by passing it over the substance known as chloride of calcium, which readily takes up moisture. Hence, pack a glass tube *e* with some lumps of chloride of calcium, and supporting it on a ring of a retort-stand or by other means, connect it with a bottle *f* in which Inflammable Air is prepared. Attach to the other end of the tube a glass tube *g* bent as shown in the figure, so that the

[1] This form of the experiment is due to Mr. C. M. Stuart; *see Nat.* liii. 347.

end *h* is a little to one side of the bottom of the retort. The end *h* of the tube must be narrowed.

As soon as all atmospheric air has been swept out, light the Inflammable Air at *h*. The flame impinges against the cold surface of the retort, and a dew is seen to be formed. Place a small beaker immediately below the retort. Presently a drop of the dew collects and falls into the beaker.

We must obtain sufficient of the dew to enable us to examine its nature. Enough will be obtained in half an hour.

EXPT. 5. **To find whether the Dew is Water.**—The chemist has tests by which he can recognise very many substances. But water, one of the most common, is not one of the easiest to test for.

Examine the liquid as regards :—

(*a*) *Smell.*

(*b*) *Taste.*

(*c*) Powder some crystals of blue copper sulphate. Heat them gently in a porcelain crucible until they become quite white. Break the white lumps down with a glass rod. Allow them to get quite cold Add a drop of the liquid to a portion. The copper sulphate turns blue again.

(*d*) *Boiling-Point* :—Tie a small wisp of cotton wool by thread round the bulb of a thermometer. Moisten the wool with the liquid. Introduce the bulb into a dry test-tube, and hold the tube nearly level a couple of inches above a small flame. Watch the rise of the mercury. It is arrested at 100° C. and then continues to rise.

(*e*) *Freezing-Point* :—Again moisten a wisp of cotton wool round the bulb of a thermometer with the liquid and introduce it into a dry test-tube, but plunge the tube into a freezing-mixture, and watch the fall of the mercury. It is arrested at 0° C. and may then sink further : if so, the melting point may also be observed.

What do you conclude from your observations? The liquid has the properties of water. Hence, it may be concluded that it *is* **water.**[1]

Recall Experiment 2. It seems that **water must be a compound of Inflammable Air and oxygen.**

[1] Cavendish proved that under certain conditions a little nitric acid is also formed and can be found in the dew.

In order to make still more sure of this conclusion, Cavendish exploded together large quantities of Inflammable Air and oxygen, and found the dew deposited to be water.

(B) THE ANALYSIS OF WATER

We have seen how Cavendish proved in 1782 by *synthesis* that water is a compound of Inflammable Air and oxygen. In the following year Lavoisier made this conclusion certain by effecting the *analysis* of water.

The problem to be solved may be stated thus. Given that water is a compound of oxygen and Inflammable Air, how can these be separated from water? Now we know of several substances which combine with the oxygen of the air when

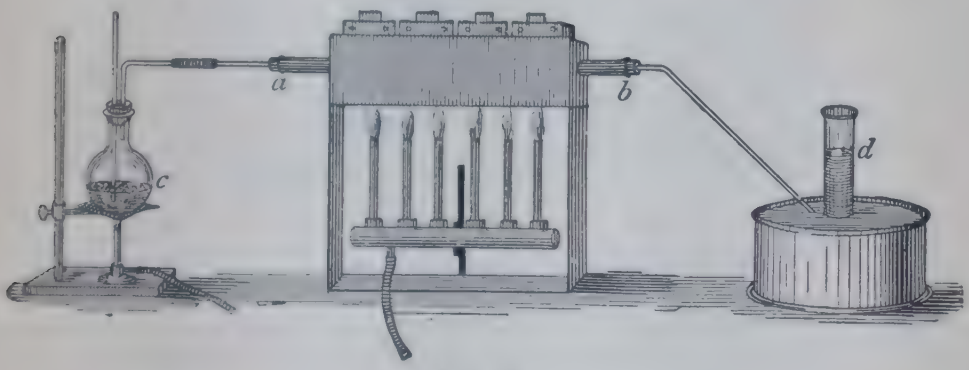

FIG. 39.

heated, for instance, iron, magnesium, and others. Is it not possible that one of these may combine with the oxygen which we are now assuming to be in water, just as in a former experiment we succeeded in removing oxygen from nitric oxide ('Nitrous Air') by passing it over heated metals? If so, Inflammable Air may be set free. Of course at the ordinary temperature most metals have little or no action on water; but suppose we try passing water vapour over red-hot iron.

EXPT. 6.[1] **To find whether Water is decomposed by Red hot Iron.**—Pack an iron gas pipe 2 feet in length with clean iron or steel tacks. Place it in the bed of a furnace (Fig. 39). Light the furnace, so that the tube is heated to a

[1] Lecture Experiment.—The nails may be cleaned from grease by heating them in the furnace whilst a current of coal gas is passed over them.

bright red heat. Attach a delivery tube to the end *b* of the iron tube, and have a pneumatic trough and gas jar at hand. Place some water in a 16 oz. flask *c*. Fit it with a cork, through which are passed a straight glass tube dipping below the surface of the water and a bent glass tube, so that it can be connected with the end *a* of the tube in the furnace. Boil the water in the flask. Then connect the flask with the iron tube, so that the steam is passed over the red-hot iron. Place an inverted gas jar *d* full of water over the end of the delivery tube. A gas bubbles up and is collected in the jar. Neglect the first jar which is filled. Fill others. Test the gas with a lighted splint. Can it be poured upward? Inflammable Air has been set free, as we

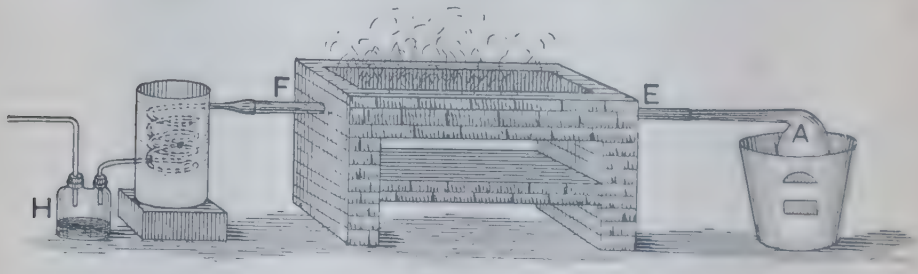

FIG. 40.

hoped. Allow the iron tube to cool, and extract some of the iron nails. Examine them.

It is of interest to note the way in which Lavoisier himself carried out this experiment.

Lavoisier's Experiment to show the Decomposition of Water by Iron.—A gun barrel E F was filled with iron filings, and placed in a furnace (Fig. 40). To the end E was luted a glass retort A containing a known weight of pure water. The end F was connected with a worm surrounded by cold water. The other end of the worm passed into a bottle H. A delivery tube was led from the flask H to a pneumatic trough. The tube EF was heated red hot. On lighting a fire under A, the water boiled and the steam passed down the gun barrel. The retort was heated until the water had entirely evaporated. Some of the steam passing through the furnace unchanged was condensed to water in the worm, and collected in H. The rest was decomposed by the iron, with the result that a gas was set free. This was collected at the trough, and proved to be

Inflammable Air. At the end of the experiment Lavoisier examined the iron in the tube. It had *gained* in weight ; it was "in the state of the black oxide, precisely like that which had been burnt in oxygen." Moreover, Lavoisier observed the quantities obtained ; thus, to quote an experiment,

Weight of water decomposed = weight of water taken in A *less* weight of water collected in H = 100 grains.

Weight of Inflammable Air produced = 15 grains.

Increase in weight of the iron = 85 grains.

This was the First Analysis of Water ever made.—Observe the conclusion to which Lavoisier came. "One hundred grains of water have been decomposed ; 85 of oxygen have combined with iron, converting it into the black oxide, and 15 grains of Inflammable Air have been liberated ; *therefore*, water is composed of oxygen and Inflammable Air in the proportion of 85 parts to 15."

Problem

Find whether steam can be decomposed by (*a*) magnesium, (*b*) zinc, (*c*) copper, with liberation of Inflammable Air, employing the apparatus shown in Fig. 41 :—

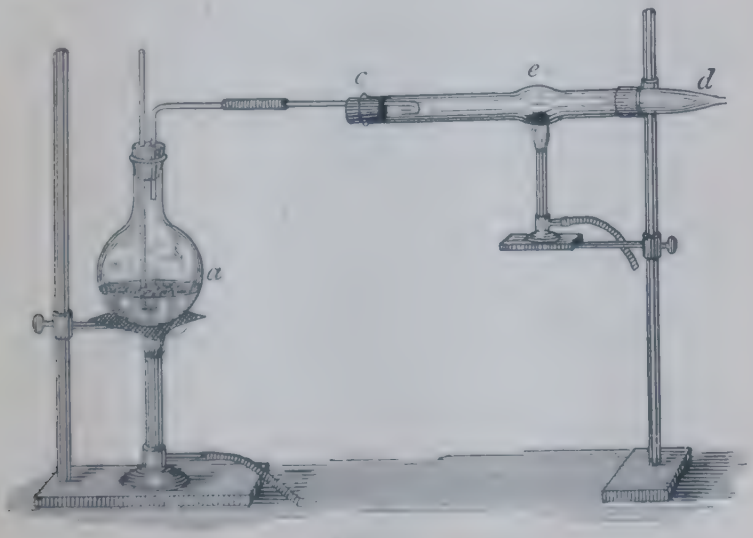

Fig. 41.

a is a flask in which water is boiled. Through the cork is passed a straight piece of glass tubing dipping below the surface of the water and a tube bent at right angles. *cd* is a tube of hard glass on which a

bulb *e* has been blown,[1] fixed in a slightly inclined position, and at the end *d* fitted with a delivery tube (not shown). The cork at *e* is connected with the flask *a* by a piece of glass tubing, one end of which is bent back on itself as shown. Place the metal in *e*. Pass steam through the apparatus, warming *cd* gently, so that the steam may not condense in it, and then heat the bulb *e*.

Find in each case (1) whether Inflammable Air can be collected from *d* over water in a test-tube; (2) whether the metal is oxidised.

A new Name for Inflammable Air.—Lavoisier, after making this analysis, at once proposed a new name for Inflammable Air. He said: "No name has appeared to us more suitable than that of **hydrogen**, that is to say, 'generative principle of water'; from *hudor*, water, and *geinomai*, I produce." This term was universally adopted, and from this point it will be used in this book.

It is to be noticed that in Lavoisier's analysis of water only *one* constituent was obtained *free*. Carlyle and Nicholson in 1800 first showed that water could be decomposed by the electric current, *both* constituents being liberated.

EXPT. 7.[2] **To show the Decomposition of Water by Electricity.**—Take a bottle with a wide mouth, and cut off the lower half horizontally, as follows:—Make a scratch on the bottle with a file. Take a pointed piece of glass rod, heat the end red hot, and touch the scratch with it. The glass cracks, and the crack can then be led right round the bottle by means of the hot glass rod, if this is reheated from time to time.[3] Choose a 2-holed rubber cork to fit the mouth, two short jets of glass (wash-bottle tips) to fit the holes in the cork, two platinum wires each 6 in. long, and two pieces of sheet platinum each 2 in. by $\frac{1}{2}$ in. Put the wires through the glass tubes and seal them in by melting the glass. Fit the glass tubes through the cork. Along the middle line of each platinum sheet prick six holes. Thread the platinum wires in and out through the sheets. Stand the bowl in a tripod stand and fill it with water (Fig. 42).

Connect the projecting platinum wires by screw clamps to the

[1] Suitable tubes are sold as *reduction tubes*.

[2] Lecture Table Experiment.

[3] Better still, buy ready made a 6 in. glass "dialyser," a kind of basin with a bottle mouth.

copper wires from a 4-celled Grove's battery[1] or three secondary cells. Nothing is noticed. Add a few drops of dilute sulphuric acid to the water. There is a rapid effervescence from the surface of the platinum sheets. But this soon slackens. Small bubbles have covered the platinum, preventing contact with the water and hindering further action. This could have been avoided by dotting the platinum sheets all over with tiny holes

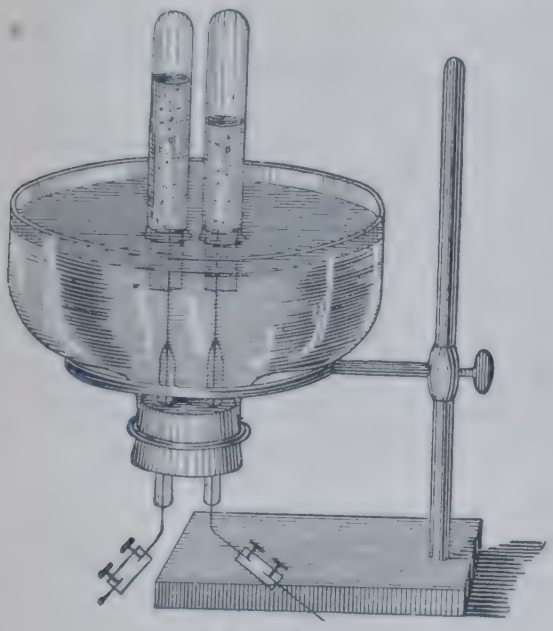

FIG. 42.

pricked by a sharp needle. From the roughened surfaces the bubbles then detach themselves easily.

From what we now know of the composition of water it seems possible that the bubbles may be oxygen or hydrogen. Fill a test-tube with water, invert it over both the platinum sheets, and collect the gas. When half full close the end with the thumb, remove the tube, and apply a light. A sharp explosion rings out. The gas is, therefore, probably a mixture of oxygen and hydrogen.

Is the same explosive gas given off from each platinum sheet? Fill two test tubes with water and invert one over each sheet. Again the current of gas slackens. But raise the tubes so that

[1] *See* Appendix for Management of Grove's Battery.

the platinum sheets are no longer shielded from each other. The bubbles come freely as before, and collect in the tubes. Gas collects much faster in one tube—about twice as fast as in the other. Find with which metal in the battery this is connected. Test the gas in each tube with a burning splinter of wood. The gas which collects slowly quickens the burning of the wood. It is *oxygen*. The other gas burns with the pale blue flame of *hydrogen*.

This experiment seems to afford additional evidence that water consists of oxygen and hydrogen, though the student may regard the small amount of sulphuric acid added to the water with suspicion until its part in the change has been clearly proved. In order to complete this experiment it is also clearly desirable to show that nothing else but hydrogen and oxygen results from the decomposition of water by electricity. This was shown after long inquiry to be the case by Sir Humphry Davy.

Water, therefore, is a compound of oxygen with hydrogen, and may be regarded as an oxide of hydrogen.

Problems

1. Find whether dry carbonic acid gas is decomposed by red-hot iron: pass any gas which issues from the tube through a strong solution of caustic soda (to remove undecomposed carbonic acid), and collect over water. Examine carefully any gas collected.

2. Find whether dry carbonic acid gas can be decomposed by passing it through a red-hot glass tube containing pieces of charcoal: wash and collect as in No. 1.

3. Find whether steam is decomposed by passing it through a red-hot tube containing pieces of charcoal. Take care not to breathe any gas which issues from the tube.

CHAPTER X

THE COMPOSITION OF WATER BY VOLUME AND BY WEIGHT

We have already established beyond doubt the fact that water is a compound of hydrogen and oxygen, and of nothing else. Moreover, we already have some evidence as to the relative volumes and the relative weights of hydrogen and oxygen which combine together when water is formed, or are liberated when water is decomposed. This question of the quantitative composition of water may now be considered more fully.

A. THE COMPOSITION OF WATER BY VOLUME

There are two distinct lines of inquiry to be followed, just as when we were investigating the composition of air, and that of carbonic acid gas.

1. Analysis.—We have already learnt that when water is decomposed by the electric current, two volumes of hydrogen are liberated for every one volume of oxygen. So far then as the analysis of water can afford evidence, the answer is definite that water consists of two volumes of hydrogen combined with one volume of oxygen.

2. Synthesis.—*Priestley* found that when hydrogen was exploded with oxygen in a bottle, the loudest explosion was obtained when about two-thirds of the bottle was filled with hydrogen and the rest with oxygen.

Cavendish showed in 1781 that 1000 volumes of air always contain nearly 210 volumes of oxygen. He abstracted the

oxygen with nitric oxide, just as Dalton had done. He arrived at the conclusion that common air contains 20·83 per cent. of oxygen by volume. In the same year he went on to explode common air with different proportions of hydrogen (*see* p. 91). He found that it was necessary to add 423 volumes of hydrogen to 1000 volumes of common air in order to burn the whole of the oxygen in the air. Almost all the hydrogen, too, was consumed. His experiment, then, proved that the water formed consisted of

$$\left\{\begin{array}{l}210 \text{ volumes of oxygen.}\\420 \text{ volumes of hydrogen.}\end{array}\right.$$

In other words, one volume of oxygen and two volumes of hydrogen united together forming water. Although the experiments of Cavendish justified this conclusion, he did not state it in this form.

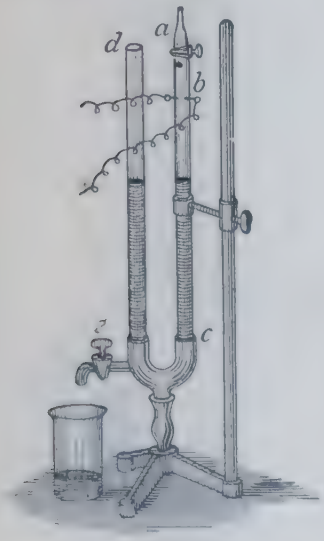

FIG. 43.

Further experiments carried out by the French chemist Gay Lussac and by Von Humboldt were necessary to lead chemists to regard the composition of water by volume established.

***EXPT. 1.[1] To find the Volume of Hydrogen combined with One Volume of Oxygen in forming Water.**

(1) *Obtain a U-shaped "eudiometer."* This is a U-shaped tube, one limb of which is closed by a glass tap *a* (Fig. 43). Two platinum wires are fused into the glass near the top of the limb *b*, and the limb may be graduated or plain. A tap *c* is fixed to the tube near the bottom of the other limb. Close the tap *c* and open the tap *a*. Pour mercury cautiously down the limb *dc*. When the mercury has risen to *a* and expelled all the air, close the tap *a*. Open the tap *c* and run out the mercury slowly. The level of the mercury remains unchanged in the limb *ab*, but it sinks gradually in the limb *da*. As soon as the

[1] Lecture Table Experiment.

level of the mercury has almost sunk to the tap *c*, close the tap *e*. The eudiometer is then ready for use.

(2) *To prepare the oxygen.*—Heat some powdered potassium chlorate in a small glass retort, and after allowing the oxygen evolved at first to escape, collect the rest in a small gas-holder over water ready for use. The "oxygen" supplied for use with magic lanterns is not pure enough.

(3) *To introduce the hydrogen.*—The hydrogen must be prepared in the usual way from zinc and sulphuric acid, taking great care to avoid any admixture of air. For this purpose Kipp's apparatus[1] is particularly suitable. After all air has been expelled from the apparatus and the leading tube, connect this tube with the tap *a* of the eudiometer, and at once open that tap. Hydrogen passes in, and the mercury sinks down *ac* and rises up *ed*. As soon as the level of the mercury is nearly the same in both arms, close the tap, and at once disconnect the supply of hydrogen. Make the level of the mercury the same in both limbs, either by adding mercury to the limb *de*, or by running some out of the tap *e*, according as it is necessary. Read the volume occupied by the hydrogen. In an experiment it was found to be 54 c.c.

(4) *To introduce the oxygen.*—Run some mercury out of the tap *e*, so that it stands at a lower level in the open limb than in the other. Attach a rubber tube to the gas-holder, run out a little oxygen, and then connect with *a*. Open *c*, and introduce about 10 c.c. of oxygen, and then close *a*. Re-level the mercury in both limbs, and read the volume of gas enclosed. In the same experiment the total volume was found to be 66 c.c. Then, 12 c.c. oxygen had been added.

(5) *To fire the mixed gases.*—Effect this by connecting the wires at *b* with a Ruhmkorff coil and a bichromate cell.[1] The mercury at once rises part way up *ac*. Pour some mercury down *de* so as to make its level the same again in both limbs. Read the volume occupied by the gas. In the experiment quoted it was 29·5 c.c. Then, the contraction was 66 − 29·5 or 36·5 c.c. Why did a contraction take place? Because the steam formed by the union of the oxygen with some of the hydrogen was condensed, and occupied a negligible volume.

(6) *To show that the gas left is hydrogen.* Pour some

[1] *See* Appendix.

mercury into *d*. Open the tap *a*, and apply a flame to the issuing stream of gas. It burns with a pale flame. Then it is hydrogen, and there was more than sufficient hydrogen to unite with all the oxygen.

(7) *Conclusion.* Hence, in the experiment above referred to, the volume of hydrogen combined in the water formed was

$$36{\cdot}5 - 12 \text{ or } 24{\cdot}5 \text{ c.c.}$$

Now this was almost exactly twice the volume of the oxygen with which the hydrogen combined.

Therefore, this experiment leads to the conclusion that **two volumes of hydrogen and one volume of oxygen, are combined in forming water.**[1]

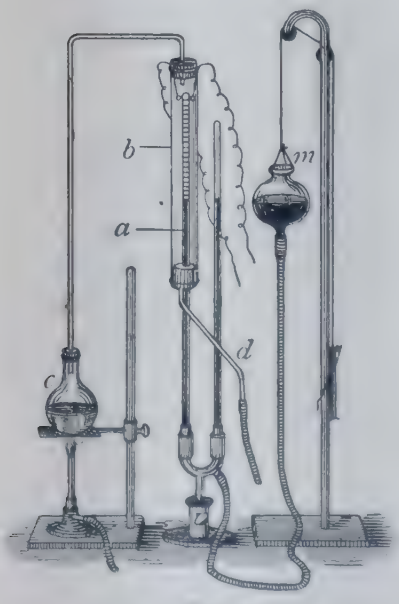

FIG. 44.

In 1808 Gay Lussac carried the investigation a stage further, and measured the volume of steam formed on the combination of hydrogen and oxygen.

*EXPT. 2.[2] **To find the Volume of steam formed when Two Volumes Hydrogen combine with One Volume Oxygen.**—Since it is desired to measure the volume of steam produced on exploding a mixture of two volumes hydrogen and one volume oxygen, it is essential to prevent the condensation of the steam. This is secured by surrounding the limb *a* of a eudiometer tube containing the mixed gases with a wider tube *b*, through which the hot vapour of boiling amyl alcohol can be passed (Fig. 44). The alcohol, which boils at 132° C., is heated in the flask *c*. The vapour after passing through the jacket is led away by the tube *d*.

The mixture of two volumes hydrogen and one volume oxygen

[1] Morley and Scott (1895) determined the ratio of the volume of hydrogen and oxygen to be 2·00269 : 1.

[2] Lecture Table Experiment.

are introduced into the eudiometer tube *a*. When the temperature of the jacket is constant at 132°, bring the mercury to the same height in both limbs. Read off exactly the volume of the mixed gases. Next reduce the pressure on the gases by lowering the reservoir of mercury *m*. Now fire the mixture by an electric spark. Raise the reservoir again until the level of the mercury is the same in the two tubes. Read the volume. It is almost exactly two-thirds of that of the mixture of hydrogen and oxygen.

Hence, **when 2 volumes hydrogen and 1 volume oxygen combine, 2 volumes of steam or gaseous water are formed.**

B. THE COMPOSITION OF WATER BY WEIGHT

(1) ANALYSIS.—Lavoisier was the first to carry out an analysis of water, by passing a known amount of water in the form of steam over red-hot iron and determining the amount of hydrogen set free (*see* page 94). Accurate results, however, could not be obtained by such a method.

(2) SYNTHESIS.—Water is produced when hydrogen is burnt or exploded with air. Then, if the water formed were weighed, and also *either* the oxygen *or* hydrogen which combined together, we should be able to arrive at its composition by weight. But it would be difficult to weigh either the oxygen or the hydrogen accurately.

Can water be synthesised in any other way? Is it possible that hydrogen can obtain the necessary oxygen from any other source than the air or oxygen gas itself? Can hydrogen obtain the oxygen from any oxide? Let us pass hydrogen over an oxide in a glass tube and try aiding the action by heat.

EXPT. 3. **To find whether Water is formed by the action of Hydrogen on Oxide of Copper.**—Dry some copper oxide by heating it gently in a basin, and arrange it in a thin layer in a porcelain boat within a hard glass tube. Place over a Ramsay burner and pass in dry hydrogen (*see* Fig. 36). Is there any action while cold? No. *After sweeping out the air* with the hydrogen, light the burner. Notice that a dew is very soon deposited on the cooler portion of the tube near *b*,

and that the surface of the oxide has become red in colour. Seemingly the copper has in part lost its oxygen and the dew is water formed by this oxygen uniting with some of the hydrogen gas. Test the dew and the red contents of the boat.

We shall now be able to find the composition by weight of water, if we can weigh the whole of the water formed in the above experiment, and also the loss of oxygen from the copper oxide.

EXPT. 4. **To find the Composition of Water by Weight.**—Take a U-tube *ab*, of the form shown in Fig. 45, with two side tubes, on one of which a bulb *a* is blown. Nearly fill it with small lumps of dry calcium chloride. Close the two

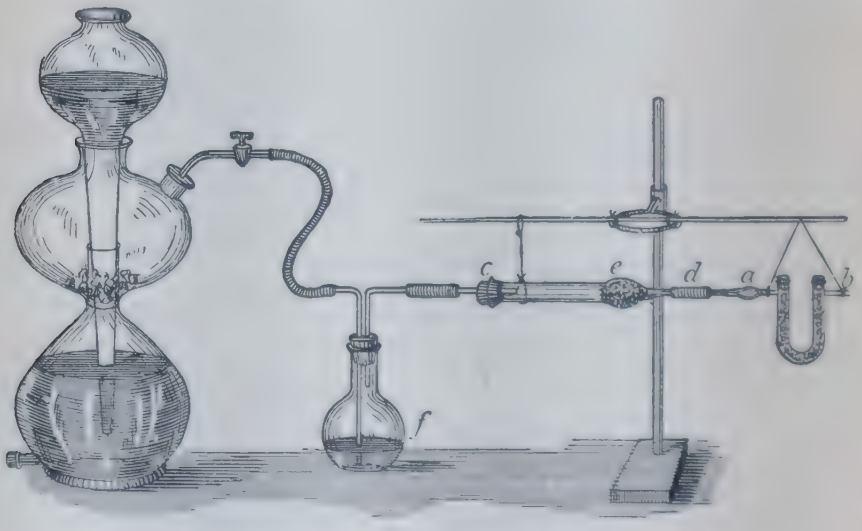

FIG. 45.

mouths with tightly fitting rubber stoppers, or better still, seal them in the blowpipe flame. Put a cap of rubber tubing closed by a small glass rod over the end of each side tube.

Dry some copper oxide (granular or in wire) by heating it strongly in a porcelain basin. Obtain a tube *cd* of hard glass on which a bulb *e* has been blown, and drawn out at one end.[1] Plug one end of the bulb loosely with a little dry asbestos. Fill the bulb with some of the copper oxide while still warm. Introduce another plug of asbestos. Close each end of the tube with stoppers or caps. Take a hydrogen Kipp apparatus.

[1] These are sold as reduction tubes

Connect it with a flask *f* containing some pure strong sulphuric acid, so that the hydrogen can be bubbled through and dried by the acid.

Weigh the tube *cd* (unstoppered) accurately. Connect it by a rubber stopper with the flask, and support it as drawn. Cause a slow current of hydrogen to pass through; it should be possible to count the bubbles as the gas passes through the acid in *f*. Weigh the U-tube (unstoppered) accurately. Fix it to the end *d* of the tube by a small piece of rubber tubing; the two glass tubes should be brought close together. Now heat the bulb *e* with a Bunsen burner, at first gently and then more strongly. The water which is formed is carried forward, and is for the most part condensed in the little bulb *a* on the U-tube. If any dew is found condensed in the hard glass tube between *e* and *d*, the tube must be gently warmed. When 1 to 2 grams water appear to have collected in the bulb *a*, gradually cease heating *e*. Allow the bulb to cool in the current of hydrogen during half an hour. Detach the Kipp and flask. Suck air through the tube *cd* and the U-tube for a few seconds. Why is this necessary? Weigh the U-tube, and also the tube *cd* accurately. Here is the result of an experiment :—

Gain in weight of U-tube, *i.e.* weight of water formed = 1·191 gm.
Loss in weight of copper oxide, *i.e.* weight of oxygen lost = 1·067 gm.
Difference, or weight of hydrogen in water = ·124 gm.

$$\text{Hence, } \frac{\text{weight of oxygen}}{\text{weight of hydrogen}} = \frac{1{\cdot}067}{{\cdot}124} = \frac{8{\cdot}6}{1}.$$

The experiment should be repeated several times. Do you obtain a similar result? The above result agrees fairly well with that which is always established by accurate experiments, viz. that **water consists of 8 parts by weight of oxygen for every 1 part by weight of hydrogen.**

Reduction. Heated copper oxide, over which hydrogen is passed, loses oxygen and metallic copper results. The copper oxide is said to be *reduced*. *Reduction*, then, is the reverse of oxidation, and consists in the removal of oxygen from a substance which contains it. Note that at the same time that the

copper oxide is reduced the hydrogen passing over it is oxidised.

In a former experiment, see p. 93, steam was passed over red-hot iron, with the result that hydrogen was liberated, and oxide of iron formed. In that case the vapour of water was reduced, while the metal was oxidised.

In both of these cases the two processes of oxidation and reduction took place simultaneously.

Problems

1. Find whether (*a*) water is produced, (*b*) the oxide is reduced to metal, by the action of hydrogen upon dry oxide of iron, red-lead, litharge, oxide of tin, manganese dioxide, oxide of magnesium, quick-lime, zinc oxide, bismuth oxide, cobalt oxide, antimony oxide.

2. Find the effect of passing a stream of coal-gas over hot copper oxide, red-lead, and oxide of iron.

3. Find by reduction the weight of oxygen combined with 1 gram of copper in the black oxide of copper.

CHAPTER XI.[1]

THE PROPERTIES OF GASES.—THE DENSITIES OF GASES

Solids, Liquids, and Gases

MANY liquids and solids, as we well know, change into gases or vapours when they are sufficiently heated. Water and iodine are readily vaporised, and ether still more so. But the vapours produced are not permanent, and when cooled are condensed or converted into liquids or solids again. We know, however, that there are other substances which are gases at the ordinary temperature. Two hundred years ago, atmospheric air was the only one known. Following the lines of discovery successfully pursued by Priestley, Scheele, Lavoisier, Black and others, we now know of oxygen, nitrogen, hydrogen, carbonic acid gas, nitric oxide and oxide of sulphur, besides common air. Whether these gases will remain gases at all temperatures is so far as we can say at the present point in our enquiry an open question. Since the vapours which can be obtained from liquids are liquefied again by cold, it seems possible, at any rate, that some of these gases might be changed into liquids at a low temperature. We cannot, however, stop to investigate this now.

We are acquainted, then, with three states of matter. But what are the essential differences between solids, liquids, and gases?

Solids.—RIGID SOLIDS.—Take a piece of steel. It is called a *solid* body. It has a particular shape of its own. Now try to

[1] *To Teachers.*—Chapters XI., XII., and XIII., on "The Properties of Gases," may if desired be postponed.

change its shape by pressing on it with a pair of pincers; it is not changed in shape perceptibly. And so, in general, a body which is not easily changed in shape when subjected to pressure in a given direction, or to a pull, or to a twist, or to any combination of pressure, pull, and twist (called in general a *stress*), is said to be a *rigid solid*.

If we subject steel to a very great stress in a given direction we find by careful measurement that a change of shape does take place, though the change is small. For instance, a piece of steel pianoforte wire can be extended $\frac{1}{86}$ of its length by a pull of 106 kilograms weight. Moreover, we find that *a given stress will always produce a definite change in shape* (called in general a *strain*). This last property belongs not only to rigid solids like steel, but to all solids. It is the characteristic property of solids.

SOFT SOLIDS.—If we take the pincers and exert upon a piece of india-rubber a pressure exactly similar to that which we exerted on the piece of steel, we find that a much greater change of shape is produced. The difference between hard and soft solids lies, therefore, in the ease with which a change of shape may be produced. But with both hard and soft solids a limited pressure produces a limited change of shape.

ELASTICITY OF SOLIDS WITH REGARD TO CHANGE OF SHAPE.—If a solid body after being subjected to a stress is relieved from the stress, it may recover either completely or partially its original shape. If it is capable of recovering its shape completely under certain conditions, it is said to be *perfectly elastic with regard to change of shape* under those conditions. For instance, if a weight of 106 kilograms be attached to a suspended steel pianoforte wire its length will be increased by $\frac{1}{86}$ of the whole; but as soon as the load is removed the wire will recover its original length. The steel is perfectly elastic under these conditions.

A solid body which does not recover its shape after subjection to a stress is said to be *plastic*. Soap, for instance, is plastic. Soft solid bodies are more often plastic than elastic. India-rubber, however, is both soft and elastic.

ELASTICITY OF SOLIDS WITH REGARD TO CHANGE OF VOLUME.—Suppose we subject a solid body, such as a piece of steel or india-rubber, to a uniform squeeze; that is to say, suppose

we exert equal pressures at the same time on all its surfaces ; we shall tend to diminish its *volume*. In the case of steel a small pressure produces no perceptible diminution in volume, and even the largest pressures we can exert produce so small a diminution in volume that we find it hard to measure it. Such bodies are said to be difficultly compressible, or by an excusable exaggeration *incompressible*. India-rubber, on the other hand, is quite easily compressible. Most solids, whether rigid like steel, or soft like india-rubber, recover their original volume, nearly or entirely, after compression (provided this has not exceeded a certain limit.) They are then said to have *elasticity of volume*. A body which recovers completely its original volume after compression under certain conditions is said to be perfectly elastic with regard to change of volume under those conditions.

Liquids.—If we melt a piece of ice in a glass beaker we obtain a certain amount of water, which we call a *liquid*. When water is at rest it is usually, as in the present case, partly in contact with a solid body, and the part in contact with the solid body takes the shape of that body. The portion not in contact with a solid body (or with another liquid) is called its *free surface*. When at rest on the earth the greater portion of this free surface is *horizontal* or perpendicular to the direction of a plumb line.

We see at once that a liquid differs considerably in properties from a solid. It has under ordinary circumstances no rigidity at all. It yields indefinitely before the slightest pressure exerted in a single given direction. If we try to squeeze a portion between our pincers it simply *flows* round them.[1] There can, therefore, be no question of elasticity with regard to change of shape.

Liquids are almost without exception difficultly compressible, or "incompressible." For example, water is compressed only $\frac{1}{11\cdot5}$ of its volume under the enormous pressure of 2000 atmospheres. They have, moreover, perfect or almost perfect elasticity of volume.

Gases.—Any one of the gases we know, oxygen for instance,

[1] *Mobile* liquids, like water, yield not only indefinitely but almost instantaneously to the slightest pressures. *Viscous* liquids like treacle yield indefinitely but slowly.

differs from a solid or a liquid in that it cannot be contained in an open vessel. If we put oxygen into a closed vessel, we find that, however big the vessel, it quickly fills it entirely and exerts a pressure on every part of the vessel. And if we were to make a hole anywhere in the vessel the oxygen would escape.

Like liquids, gases have *no rigidity*.

Unlike solids and liquids, all gases are *easily compressible*.

EXPT. 1. **To show that a Gas can change its Volume.**—Partly fill a bladder, or an india-rubber balloon with air or some other gas. Tie it up. Place it under the receiver of an air-pump, and exhaust the air (Fig. 46). The balloon swells visibly. Admit the air again. It shrinks to its former size. Gases are much more compressible than solids or liquids, and recover their original volumes completely after compression. But we shall study the compressibility of gases more closely presently.

FIG. 46.

Another difference to be noted between gases on the one hand, and solids and liquids on the other, is that gases are as a rule much *less dense* or lighter, bulk for bulk, than liquids and solids. We shall now proceed to consider the densities of gases more fully.[1]

THE DENSITIES OF GASES

Now that we are acquainted with several gases, which are permanent under ordinary conditions, it will be well to study them more closely, so as to attain a better knowledge of the gaseous state.

We know in a general way that gases differ in density or weight. A soap bubble filled with hydrogen or a balloon full of coal gas will rise in the air. Hydrogen can be poured upwards, and carbonic acid downwards. We have not, however, as yet made any exact measurements of the relative densities of gases.

[1] NOTE.—In this chapter the densities of gases are discussed apart from the influence of temperature and pressure.

Now the mass of a unit volume of any substance is called its density. Hence for gases, *Density means the mass of* 1 *litre of the gas.* For example, experiment has already shown that 1 litre of hydrogen weighs ·09 gram. Hence it is said that the density of hydrogen is ·09.

It is often convenient to compare the density of a gas with that of some other well-known gas. *Air* is frequently taken as the standard. Then the *Relative Density of a Gas means the weight of any volume of the gas divided by the weight of an equal volume of the standard gas, air.* For example, if the relative density of carbonic acid gas is stated as $1\frac{1}{2}$, it means that it is $1\frac{1}{2}$ times as heavy as an equal volume of air.

The Density of Air.—The first attempt to weigh a gas was made in the case of air, which for a long time was the only gas known. This was tried by Boyle about 1660.

EXPT. 2.[1] **An Attempt to weigh Air.**—Obtain an open gold-beater's skin balloon. Place it along with a piece of thread on one pan of a balance, and weigh it. Attach it to a bellows, inflate it, and tie up the neck with the thread. Weigh again. The weights are the same. Why is this?

Think of the balloon lying on the balance pan, open at the mouth. Air is inside it and around it and the pressure of the air is the same in all parts. Why should any passage of air in or out of the bladder change its weight? If more air is put into the balloon, it does become heavier by the weight of this air. But in air as in water the *upward force on a body immersed in a fluid is equal to the weight of fluid displaced* (*see* Vol. I, p. 61). If the balloon has a litre more of air put into it then the upward force is increased by the weight of a litre of air. So the weight of air in the balloon is always exactly balanced by the buoyancy of the air around. It is the balloon alone, not its contents, which can be weighed on the balance.

Von Guericke's Method.—Successful experiments were made about the same time by Otto von Guericke in Magdeburg. He used a large flask with a stopcock (Fig. 47). The capacity of the flask was first determined. The flask was then emptied of air with the air-pump and weighed. Lastly the air was

[1] Lecture Table Experiment.

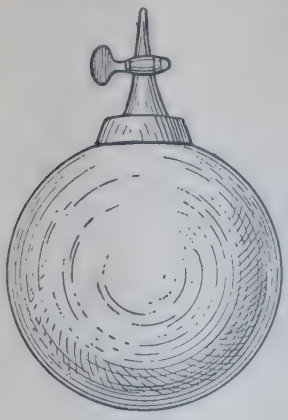

FIG. 47.

allowed to enter and the flask was weighed again. For example :—

Capacity of flask ...	750	c.c.
Weight when empty	239·0	gm.
Weight when full of air	239·9	gm.
Weight of air	·9	gm.

Then 750 c.c. of air weigh ·9 gm.

Hence, 1000 c.c. of air weigh $0{\cdot}9 \times \frac{1000}{750}$ gm.

$= 1{\cdot}2$ gm.

EXPT. 3.[1] **To find the Weight of 1 Litre of Air (Laboratory Method).** — Fit up an 8 oz. round-bottomed flask or bolt-head with a rubber cork, glass tube, rubber tube and clip, as shown in Fig. 48. Put 50 c.c. water in the flask. Support the flask on a retort-stand over wire gauze. Open the clip. Boil the water. Let it boil steadily for 5 minutes. The air which was in the flask will be swept out by steam. (What will happen if the flame is taken away from below the flask for a few seconds?) Close the clip, quickly remove the flame, thoroughly cool the flask under the tap. (Why does the water boil again?) Dry the outside of the flask. When cold weigh the flask by hanging it from the hook of the

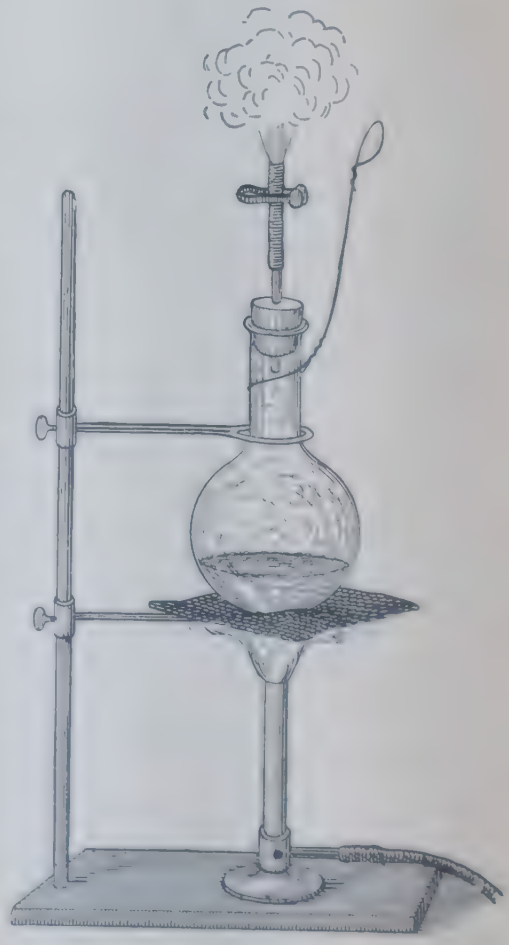

FIG. 48.

[1] Adapted from Ramsay's *Experimental Proofs of Chemical Theory*, p. 20.

balance by a wire. Open the clip. Air will hiss in. Weigh again. Find the volume of the air left in the flask, thus :—Pour the water left in the flask into a measuring jar ; find the whole volume of the flask ; subtract.

Thus,

Weight of flask without air	128·47 gm.
Weight of flask with air	128·81 gm.
Difference = weight of air	·34 gm.
Volume of water left in flask	52 c.c.
Volume of whole flask	340 c.c.
Difference = volume of air	288 c.c.
Then 288 c.c. air weigh	·34 gm.
Hence, 1000 c.c. of air weigh...	$\frac{1000}{288} \times ·34.$
	$= 1·18$ gm.

The average of many experiments shows that 1 litre of air at the ordinary temperature weighs 1·2 gram.

***Alleged Inaccuracy of the above Experiments.**—With an ordinary air-pump (*e.g.*, Tate's) it is difficult to remove more than $\frac{59}{60}$ of the air from a globe. Mercury pumps (*e.g.*, Sprengel's) will give a much greater exhaustion. Again in the last experiment it is possible that a small amount of the air may not be swept out by the steam. But there is another source of error.

As a cork floats on water or a balloon in the air, so too a glass globe is buoyed up by the air around it. Although the force is not enough to make the globe float, it does make it seem lighter. A globe which displaces a litre of air is buoyed up by a force equal to the weight of a litre of air. We have seen that this is about 1·18 gram for a litre globe. But the brass weights on the other pan are also buoyed up, though to a less extent, because they are smaller, say by a force of ·004 gram weight. Then on the whole such a glass globe must seem to weigh 1·18 − ·004 or 1·176 gram lighter than its weight in vacuo. But this error affects both weighings ; and therefore, as the results are subtracted, the errors seem to cancel each other.

Then are these corrections for the air displaced needless? The numbers 1·18 and ·004 depend on the weights of air displaced. Now the weight of the air depends, as we shall presently find, upon its temperature and pressure, and these

may change during the course of an experiment. In that case *different* corrections may have to be made out and applied to each separate weighing.

***Regnault's Method.**—This method enables the observer to conduct the experiment so that no corrections need be applied, however much temperature and pressure may vary during the experiment.

Two glass globes are required of as nearly the same capacity as possible. They are suspended by wires at the two ends of the beam of a balance without removing the pans, and they are exactly counterpoised by adding a small piece of glass tubing to the lighter globe (Fig. 49). The left-hand globe is then exhausted

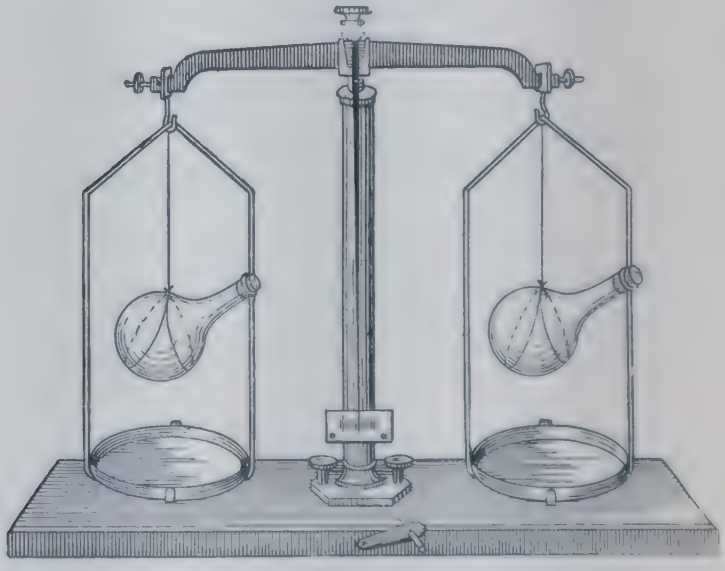

FIG. 49.

as completely as possible by an air-pump, then closed or sealed up, and replaced above the left-hand pan. To restore equilibrium the necessary weights are added to the left-hand pan. The capacity of the left-hand globe must also be determined. For example :—

Capacity of globe...	300 c.c.
Counterpoise added to empty globe ...	·363 gm.
If 300 c.c. of air weigh	·363 gm.
1000 c.c. of air weigh	$\frac{363}{300}$ gm.
	= 1·21 gm

In an experiment conducted according to this method, therefore, by employing 2 globes of equal capacity, any change in the buoyancy of the air affects both ends of the balance beam to the same extent. For the volume of air displaced by the weights used in the second weighing is so small as to be negligible.

This method may be applied to find the density of any gas. We shall, however, proceed to show how the density of carbonic acid may be determined by another method. We know that carbonic acid gas is heavier than air. What, then, does a litre of carbonic acid weigh?

EXPT. 4. **To find the Weight of 1 Litre of Carbonic Acid Gas.**—Carefully clean and dry a 16 oz. flask. Fit it with a good cork. Weigh it accurately. Remove the cork. Pass the gas from a generator containing marble and hydrochloric acid through a drying bottle containing strong sulphuric acid (to dry the gas) and into the flask, filling it by downward displacement. Test with a burning splint from time to time, and when the flask is full remove the delivery tube. Cork immediately. Weigh again. Then invert the flask in a trough containing dilute (three per cent.) caustic soda solution. Remove the cork below the surface of the solution, and agitate mouth downwards until all the carbonic acid has been absorbed. Cork it again. The size of a small air bubble remaining shows how nearly all the air is removed. Remove the flask and ascertain the volume of caustic soda in the vessel by pouring it into a measuring jar.

In an experiment,

Weight of flask filled with air ... ... = 66·855 gm.
,, ,, carbonic acid = 67·173 gm.
difference ... ... = 0·318 gm.
Volume of carbonic acid... ... ... = 497 c.c.

Now, weight of 497 c.c. air ... ... $= 1{\cdot}2 \times \frac{497}{1000}$ or 0·596 gm.

Hence, weight of 497 c.c. carbonic acid = 0·596 + 0·318
= 0·914 gm.

And, weight of 1 litre of carbonic acid gas ... .. $= 0{\cdot}914 \times \frac{1000}{497}$
= **1·84** gm.

Do your results confirm this?

Summary.—The most accurate measurements of gaseous densities have been made by Regnault's method. Recently in the hands of Lord Rayleigh, results of extraordinary accuracy have been obtained.[1] For example, the density of hydrogen is determined as ·09001, or very nearly ·09. Hydrogen is found to be very much lighter than all the common gases, and consequently it is very convenient to refer their densities to hydrogen as a standard. In that case their densities relative to hydrogen can be stated very approximately as whole numbers, as in the third column of the following table :—

	Density or Weight of 1 Litre at 0° C.	Relative Density Air = 1.	Relative Density Hydrogen = 1.
Hydrogen	·09	·0696	1
Nitrogen	1·26	·972	14
Air	1·29	1·000	14 +
Nitric oxide...	1·35	1·036	15
Oxygen	1·44	1·105	16
Carbonic acid	1·98	1·529	22
Sulphur dioxide	2·88	2·247	32

FACTS WORTH REMEMBERING :—At 0° C. :—

1 litre of air weighs 1·29 gram.
1 litre of hydrogen weighs 0·09 gram.
1 gram of hydrogen fills $\frac{1}{\cdot 09}$ or 11·11 litres.

PROBLEMS

1. Use the apparatus employed in Experiment 3, but without water inside. Weigh it dry and full of air. Suck out some of the air. Weigh again. Find the volume of air removed. Hence calculate the weight of 1 litre of air.

2. Devise a method of using the same apparatus to find the weight of 1 litre of coal gas.

[1] One splendid result of Rayleigh's accurate weighings has been his discovery (1894), in connection with Professor Ramsay, of another gas in the atmosphere, which is present in extremely small quantity. This gas has been named *argon*.

Examples I

The Densities of Gases

1. A flask weighs 130 grams when full of air, and 129·84 grams when some of the air has been sucked out. When opened under water 125 c.c. of water enter. Find the weight of a litre of air.

2. A flask of 750 c.c. capacity weighs ·9 gram more when full of air at 29 inches mercury pressure than when full of air at 2 inches pressure. Find the weight of a litre of air at 30 inches pressure.

3. If a litre of air weighs 1·29 gram, find how many c.c. 1 gram of air will fill.

4. Find the weight of 1 cu. ft. of air in ounces; assuming 1 linear foot = 30 cm. long, 1 oz. = 28 grams, 1 litre air = 1·26 gram.

5.[1] A flask of 250 c.c. volume weighs 125 grams when full of air. What will it weigh when empty?

6.[1] A closed flask of 10 litres external volume is balanced by a brass kilogram weight. Find the weight of the flask in vacuo. Relative density of brass = 8.

7.[1] If 250 c.c. of a gas weigh ·09 gram more than 250 c.c. of air, find the weight of 1 litre of the gas.

8.[1] If coal gas is half as heavy as air, find the lifting power of a balloon of 100 litres capacity.

9.[1] If 750 c.c. of a gas weigh ·39 gram less than 750 c.c. of air, find the weight of 1 litre of the gas.

[1] In questions 5 to 9 assume that 1 litre of air weighs 1·2 gram.

CHAPTER XII

THE PROPERTIES OF GASES.—BOYLE'S LAW

WE have already satisfied ourselves that the air, though invisible, has weight. About 240 years ago the Honourable Robert Boyle working in Oxford perfected the air-pump or "pneumatical engine," as he called it, and proceeded to make many interesting experiments. He was soon persuaded that "there is a spring or elastical power in the air we live in." He made the experiment described on p. 110, and saw in the expansion of the bladder as the air was gradually removed from the receiver an evidence of the pressure or force of the "spring" of the air inside the bladder.

Again, he fitted a barometer into the receiver of an air-pump and found to his delight that as the air pressure upon the mercury in the cistern was exhausted, the space above the mercury in the barometer tube increased at the same time, and the mercury fell.

This experiment suggested to him that it was worth while to investigate what relation there was between the spring or pressure of the air in the barometer tube and the volume it occupied, for, said he, "it is probable that the discovery would not be unuseful." He finally solved this question successfully in a way which we can readily repeat.

EXPT. 1. **To find the Relation between the Volume of Air and the Pressure it exerts.** Take an ordinary glass tube, about 110 cm. long, and seal one end *a* in a flame. Bend the tube at a point about 20 cm. from the sealed end, until the two limbs are parallel to one another. Cut a groove in a cork disc. Draw out the tube of a thistle funnel so that it can be

inserted in the glass tube at *c*. Support the tube by means of a clamp upon the cork. (Fig. 50.)

Now pour in enough mercury to fill the bend of the tube. Take the tube out of the clamp, and bring the mercury to the same level in both limbs of the tube by inclining the tube frequently, and so permitting air to pass from one limb to the other. Fix the tube in the clamp again. The enclosed air is now at the pressure of the atmosphere outside. Mark the level of the mercury in the closed limb at *b* with a small label.

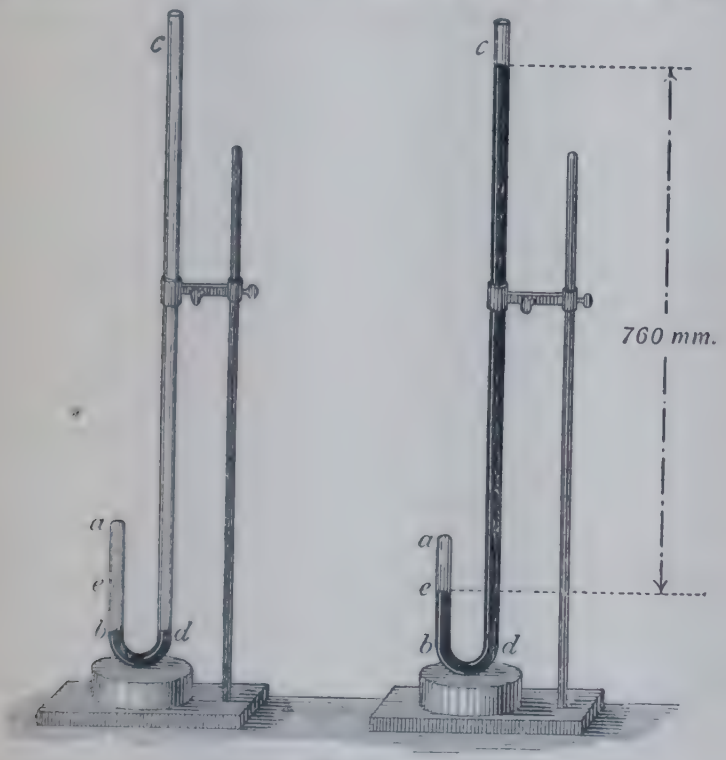

FIG. 50.

Measure the length *ab* with a metre scale. Gum a small label at a level *e*, halfway between *a* and *b*.

Then carefully pour mercury into *cd* through the funnel. The volume of the air in *ab* is diminished, for the level of the mercury rises gradually up *ba*. Proceed in this way until the mercury has risen exactly half-way up *ba*. Now observe that the mercury stands at a much higher level in the limb *cd*. Measure the *difference* in the level of the mercury in the two limbs with the metre scale. Is it not about 760 mm.?

Therefore, the pressure of the enclosed air is now balanced by that of the atmosphere and that due to 760 mm. of mercury. But the pressure of the atmosphere is equal to that of 760 mm. of mercury. Therefore the pressure of the enclosed gas is *double* the atmospheric pressure, while the gas occupies *half* the original volume.

Boyle's account (1662) of his own experiment is worth reading, and after performing the above experiment it will be easier to follow the rather involved sentences in which it is described.[1]

"We took then a long glass tube, which, by a dexterous hand and the help of a lamp, was in such a manner crooked at the bottom, that the part turned up was almost parallel to the rest of the tube, and the orifice of this shorter leg of the syphon (if I may so call the whole instrument) being hermetically sealed, the length of it was divided into inches (each of which was subdivided into eight parts) by a streight list of paper, which containing those divisions, was carefully pasted all along it. Then putting in as much quicksilver as served to fill the arch or bended part of the syphon, that the mercury standing in a level might reach in the one leg to the bottom of the divided paper, and just to the same height or horizontal line in the other; we took care by frequently inclining the tube, so that the air might freely pass from one leg into the other by the sides of the mercury (we took, I say, care), that the air at last included in the shorter cylinder should be of the same laxity with the rest of the air about it. This done we began to pour quicksilver into the longer leg of the syphon, which by its weight pressing up that in the shorter leg did by degrees streighten the included air; and continuing this pouring in of quicksilver till the air in the shorter leg was by condensation reduced to take up but half the space it possessed (I say, possessed, not filled) before: we cast our eyes upon the longer leg of the glass, on which was likewise pasted a list of paper carefully divided into inches and parts, and we observed not without delight and satisfaction that the quicksilver in that longer part of that tube was 29 inches higher than the other. . . . . Here the same air being brought to a degree of density about twice as great as that it had before, obtains a spring twice as strong as formerly."

[1] Boyle's Complete Works, vol. 1, p. 100.

Problems

Making use of the U-tube used in Experiment 1, find the change in pressure which occurs when —

(*a*) The volume is reduced to three-quarters of the original volume.
(*b*) The volume is reduced to two-thirds of the original volume.
(Do not place the tube in direct sunlight, nor handle it.)

The result of these and other experiments may be summed up in the following statement, which in honour of Boyle is usually known as **Boyle's Law.** —**When the temperature remains constant, the volume occupied by a definite mass of air varies inversely as its pressure,** *e.g.* if the pressure is increased 2, 3, 4, 5 times, the volume will be reduced to $\frac{1}{2}$, $\frac{1}{3}$, $\frac{1}{4}$, $\frac{1}{5}$, of the original volume.[1]

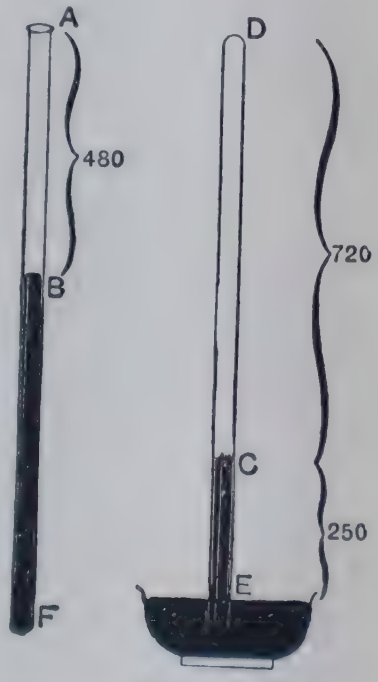

FIG. 51.

The apparatus we have used is suitable for studying the change in the volume of air when the pressure is increased above that of the atmosphere. It is not suitable for finding whether Boyle's law applies also when the pressure is reduced ; yet it is most desirable to learn whether this is so or not.

EXPT. 2. **To test Boyle's Law for Pressures less than that of the Atmosphere.**—Take a straight tube A F (Fig. 51), about 3 feet long, sealed at one end, like the one used in Vol. I., p. 67, Expt. 7. Fill rather more than half the tube with mercury. Measure the length A B of the column of air still in the tube, say 480 mm. Hold the thumb firmly over the open end A. Invert the tube in a basin of mercury. the air rises to the top. Remove the thumb. The mercury drops and the air expands. Measure the length D C of the air

[1] Boyle must share the honour of arriving at this important generalisation with Mariotte, whose treatise *De la Nature de l'Air*, was published in Paris in 1676.

space now, say 720 mm. Also measure the length C E of the mercury column, say 250 mm. Then,

the volume of the air is $\frac{720}{480}$ or $\frac{3}{2}$ of what it was.

What is now the pressure of the air in D C? The pressure of the air in A B was that of the atmosphere. Read the barometer, say 750 mm. Now,

pressure of the air in D C + pressure of the mercury column C E } balance the external pressure of the atmosphere.

Hence,

Pressure of the air in D C = 750 − 250 or 500 mm. of mercury.

Hence, the pressure of the air is $= \frac{500}{750}$ or $\frac{2}{3}$ of what it was.

Now $\frac{3}{2}$ is the inverse ratio of $\frac{2}{3}$, and consequently the volume of the air has changed in the inverse ratio of the change in the pressure.

FIG. 52.

Boyle was acquainted with but one gas—viz., atmospheric air. We now know other gases, and therefore it is important to try whether Boyle's law is true for them also.

EXPT. 3. **To find whether Boyle's Law is true for other Gases, for example, Carbonic Acid Gas.**[1]—Fit a small strong wide-necked bottle F with a 3-holed stopper (Fig. 52). Introduce through two of the holes two glass tubes A and B each one metre long, which are drawn out near their upper ends. Fix through the third hole a tube C open at both ends. Obtain also a long glass rod which can slide freely within C. Fill the bottle with mercury. Fit the cork firmly in the neck, and bind it down with wire. Place the rod in C.

Pour mercury into the tube C until the two other tubes are full of mercury. Connect A to a generator of carbonic acid gas. Leave B open to the air. Now raise the rod slowly. The mercury falls in A and B, and carbonic acid is drawn into A, and air into B. Manipulate the supply of carbonic acid so that the

[1] Lecture Table Experiment.

mercury levels E and D are about halfway down the tubes A and B, and the same in each. Then seal the tubes at A and B.

Now lower the plunger in C. The level of the mecury in C is raised, and therefore the pressure upon the gases in A and B is increased, and to the same extent. Notice that the two surfaces E and D move up together, and are always on the same level. This shows that under the same pressure the gases are equally compressible. Also raise the plunger. The surfaces E and D sink together.

It is evident that under changing pressure the volume occupied by the carbonic acid changes in precisely the same way as that of the air. We are, therefore, justified in concluding that Boyle's law applies to carbonic acid as well as to air. In the same way it can be proved that the law is equally applicable to any of the other permanent gases with which we are acquainted.

We may therefore re-state Boyle's law thus :—

When the temperature remains constant the volume occupied by a definite mass of any gas varies inversely as its pressure.

Our experiments have only been made over a small range of pressure. We have therefore no right to assume that if the pressure on a given mass of gas is increased 100 times, its volume will be reduced exactly 100 times. All we can rightly conclude is that Boyle's law holds for gases within the range of pressure covered by our experiments and with the degree of accuracy of measurement possible in our observations.

Problem

Fit up the apparatus shown in Fig. 53, using a litre flask, rubber stopper, glass T-piece, rubber tubing, two clips, glass tubes, and basin of mercury. Open the clips. Suck at A. Mercury rises up B. Close the clips. Find whether the air in C has expanded according to Boyle's law.[1]

Mathematical Expression of Boyle's Law.—Let a certain mass of gas have a volume V_1 and a pressure P_1. Let its volume be changed to V_2 when its pressure becomes P_2.

[1] This problem is taken from Ramsay's *Experimental Proofs of Chemical Theory*, p. 23, where it is worked out in detail.

Then by Boyle's Law,

$$\frac{V_1}{V_2} = \frac{P_2}{P_1} \quad \ldots \quad \ldots \quad \ldots \quad \ldots \quad \ldots \quad (1).$$

or

$$V_1P_1 = V_2P_2 \ldots \quad \ldots \quad \ldots \quad \ldots \quad \ldots \quad (2).$$

Similarly, if it be made to take other volumes V_3, V_4, &c.,

$$V_1P_1 = V_3P_3 = V_4P_4, \text{ \&c.}$$

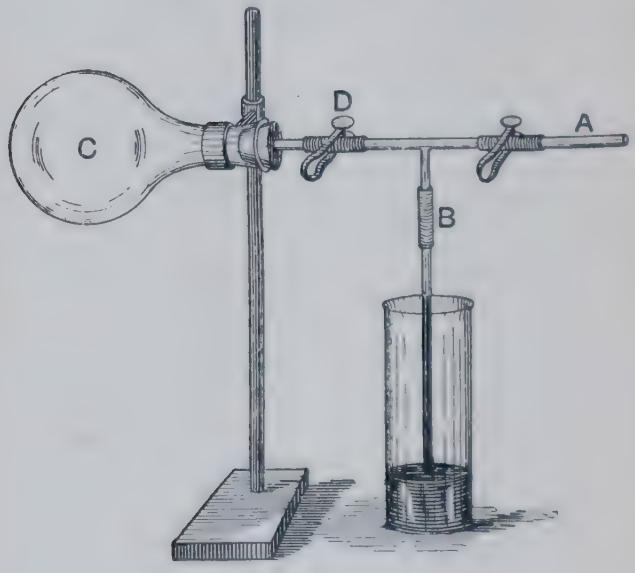

FIG. 53.

The product of pressure and volume, therefore, remains constant, or

$$\mathbf{PV} = \mathbf{C} \text{ (a constant)} \quad \ldots \quad \ldots \quad \ldots \quad (3).$$

From (2),

$$\frac{V_2P_2}{V_1P_1} = 1 \quad \ldots \quad \ldots \quad \ldots \quad \ldots \quad \ldots \quad (4).$$

This form of the equation is useful in testing the results of experiments.

EXAMPLE.—In an experiment $V_1 = 338{\cdot}5$ c.c., $P_1 = 595{\cdot}2$ mm., $V_2 = 274{\cdot}5$ c.c , $P_2 = 736{\cdot}2$ mm.

Then,

$$\frac{V_2P_2}{V_1P_1} = \frac{2745 \times 7362}{3385 \times 5952} = 1{\cdot}003+$$

Hence the above data agree very nearly with Boyle's Law.

Again, from equation (2)

$$V_2 = V_1 \times \frac{P_1}{P_2} \quad \ldots \ldots \ldots \ldots \quad (5).$$

This equation is useful in calculations like the following :—

EXAMPLE.—What volume will be occupied by 250 c.c. of air at 720 mm. when the pressure changes to 750 mm. ?

Here,

$$V_1 = 250 \text{ c.c.}, \quad P_1 = 720 \text{ mm.}, \quad P_2 = 750 \text{ mm.}$$

Hence,

$$V_2 = 250 \times \frac{720}{750} \text{ c.c.}$$

$$= 240 \text{ c.c.}$$

The Measurement of Gases collected over Water or Mercury.—When the volume of a gas standing over water is measured, the liquid levels should be the same within and without the tube. This can generally be arranged for. But if not, a calculation must be made.

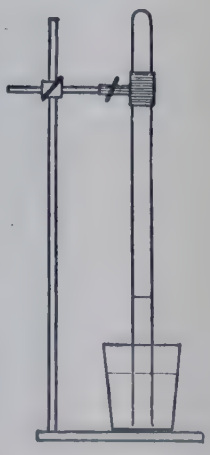

FIG. 54.

EXAMPLE.—100 c.c. of air are collected in a tube over water (Fig. 54), when the barometer is at 750 mm., the water standing 27 cm. higher inside the tube than without. What volume would the air occupy at the atmospheric pressure 750 mm. ?

Now, since the relative density of mercury is $13\frac{1}{2}$, 27 cm. of water produce the same pressure as

$$\frac{270}{13\frac{1}{2}} \text{ mm. or 20 mm. of mercury.}$$

Then, the external air pressure of 750 mm. is balanced by the internal air pressure, together with that of 20 mm. of mercury.

Hence, the internal air pressure $= 750 - 20$ mm.

$$= 730 \text{ mm.}$$

Here,

$$V_1 = 100 \text{ c.c.}, \quad P_1 = 730 \text{ mm.}, \quad P_2 = 750 \text{ mm.}$$

Hence,

$$V_2 = 100 \times \frac{730}{750}$$

$$= 97{\cdot}3 \text{ c.c.}$$

Graphic Expression of Boyle's Law. Draw two lines OX and OY at right angles (Fig. 55). Measure distances in any units, say inches, along each. Let distances along OX represent volumes, and distances along OY pressures.

The point A is distant 1 inch from OX and 1 inch from OY, and may be taken to represent a given mass of some gas at unit pressure and unit volume. Similarly, the point B represents the same mass of gas at double volume and half the pressure. Again, C represents the gas at half the volume and double the pressure. Notice that in each case the rectangle enclosed by the dotted lines, representing volumes and pressures, has an area of 1 square inch. The product, then, of pressure and volume is represented by a constant area (cf. equation 3, p. 124).

Find other points representing the same mass of gas when its volume is $\frac{2}{3}$, $\frac{3}{4}$, $1\frac{1}{2}$, &c. Notice that these points lie on a curve

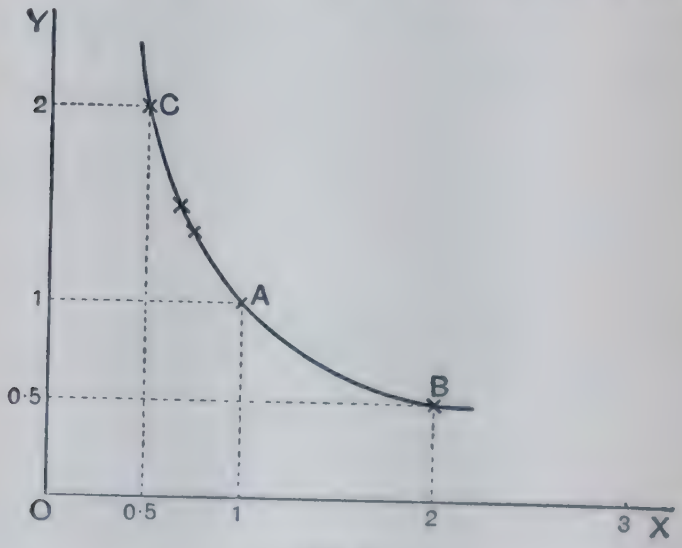

FIG. 55.

Draw the curve C A B. This line C A B is the graphic representation of Boyle's Law. The curve is what mathematicians call a *rectangular hyperbola.*

Obtain some paper ruled in squares, say ten squares to the inch, and plot all the results obtained in Experiments 1 and 2. Can a smooth curve be drawn through them?

The use of the word "Law."—"Words," wrote the Duke of Argyll,[1] "which should be the servants of Thought are too often its masters; and there are very few words which are used more ambiguously than the word 'Law.' . . . . In its

[1] *The Reign of Law*, by the late Duke of Argyll, p. 64.

primary signification a 'law' is the authoritative expression of human Will enforced by Power."

But as applied by Science to natural phenomena the word "law" has a less dignified sense; it is a brief description of observed facts. Thus if the pressure on a given mass of air at a constant temperature is doubled the volume is halved; if the pressure is trebled the volume is reduced to one-third. And a similar experience is found with hydrogen and other gases. These observed similarities are summed up in the brief description, that when the temperature remains constant the volume occupied by a given mass of gas varies inversely as its pressure, and this description is called *Boyle's 'Law.'* But as has already been stated our experiments have only been made over a small range of pressure, and it is a misuse of language to say that Boyle's Law "must" always be "obeyed." Boyle's Law is not a command given to gases which they are required to observe. We have no right to conclude that if the pressure on a given mass of gas is increased 1000 times its volume will be reduced 1000 times. It may or may not be the case; we simply do not know until it is determined by experiment.

The Laws of Science do not "govern" the universe, nor are they analogous to legal enactments. "They are mere descriptions, often tentative, often erroneous, of similarities which we believe we have observed."[1]

Examples II

Boyle's Law

1. Given $V_1 = 99$ c.c., $P_1 = 750$ mm., $V_2 = 101$ c.c., find P_2.

2. 1,440 c.c. of air were enclosed in a bell-jar over water on a day when the barometer showed 750 mm. What volume would the air occupy if the barometer had fallen to 720 mm.?

3. 100 c.c. of air were measured at 750 mm. The volume afterwards changed to 101 c.c. Find the new pressure.

4. Find the capacity in cubic inches of a steel oxygen cylinder which contains under a pressure of 1,800 lbs. per sq. in. enough oxygen to fill 20 cu. ft. under a pressure of 15 lbs. per sq. in. (assuming Boyle's Law to hold at the high pressure).

[1] Prof. J. H. Poynting. See *Nature*, vol. lx. p. 470.

5. Find the pressure in lbs. per sq. in. required to force 40 cu. ft. of coal gas at 15 lbs. per sq. in. into a space of $\frac{1}{2}$ cu. ft.

6. By how much must the pressure on 1,000 c.c. of air at 760 mm. be increased to make the volume 50 c.c. less?

7. What change will there be in the volume of 1,000 c.c. of air at 30 ins. of mercury pressure when the barometer falls to $28\frac{3}{4}$ ins.?

8. In a Boyle's Law Experiment with a syphon tube there was originally 6 ins. length of air in the shorter limb. The barometer stood at 30 ins. Mercury was poured into the open limb until it was 15 ins. higher there. Find the length of air in the inner tube (*see* Fig. 50, p. 119).

9. In another similar experiment, 12 ins. length of air was enclosed in the shorter limb. What height of mercury would compress the air to 4 ins. length?

10. Two flasks of 500 c.c. and 1,000 c.c. are connected by a stop-cock. The 1,000 c.c. flask is vacuous. The 500 c.c. flask contains air at 750 mm. pressure. Find the pressure after opening the tap.

11. Are the following figures in agreement with Boyle's Law? Find $\frac{V_2P_2}{V_1P_1}$ to 3 places of decimals.

$V_1 = 250$ cc., $P_1 = 760$ mm., $V_2 = 253$ c.c., $P_2 = 750$ mm.

12. $V_1 = 240$ cu. ins., $P_1 = 29$ ins., $V_2 = 232$ cu. ins., $P_2 = 30$ ins.

CHAPTER XIII

THE PROPERTIES OF GASES. THE EXPANSION OF GASES BY HEAT

It has already been shown (Vol. I., p. 41) that air expands when heated, and contracts on cooling. We have, however, made no endeavour to measure the amount of this expansion, nor have we tried to compare the expansions of different gases.

John Dalton made, in 1801, in Manchester, the first really important research on this subject, and shortly afterwards *Charles* in Paris arrived quite independently at the same conclusions as Dalton.

Suppose we study in the first instance the expansion of air when heated.

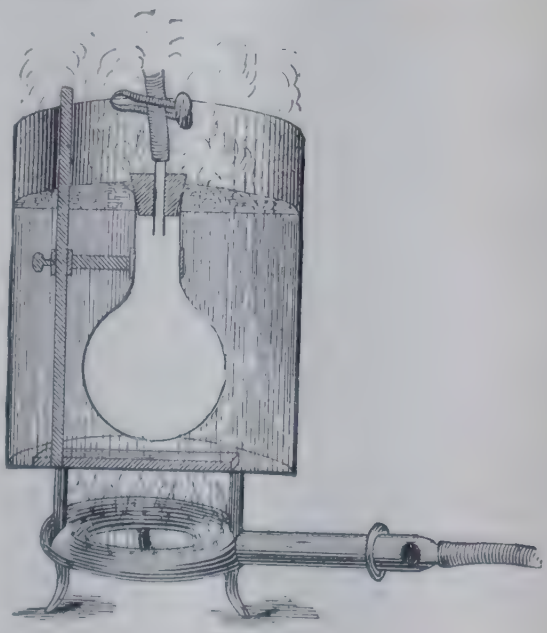

Fig. 56.

Expt. 1.[1] **To find the Expansion of 1 Litre of Air when heated 1° C.** Dry thoroughly an 8 oz. round-bottomed flask or bolt head. Choose a closely-fitting one-holed rubber stopper. Fit it with a short glass tube, rubber tube and clip[2] as shown in Fig. 56. The tube must fit very tightly into the

[1] Adapted from Ramsay's *Experimental Proofs*, p. 21.

[2] A stopper of glass rod may be used in place of the clip.

stopper. Place the stopper firmly in the flask. Fix the flask in a small retort-stand, which can be placed within a cylindrical tin vessel. Half fill the tin with water, and heat it over a rose-burner. Open the clip on the flask. As soon as the water begins to boil, place the stand with the flask in the tin so that the flask is immersed in the water up to the stopper. After a minute or two, close the clip. Heat the flask thus for at least 5 minutes, but open the clip for a moment from time to time. Take the temperature of the water. Then tighten the clip, and remove the flask. Cool the flask. Invert the flask in cold water, and open the clip under the water. Water enters the flask. Why? Shake. This will ensure the cooling of the air and flask to the temperature of the water. Lower the flask till the level of the water is the same within and without. Then close the clip under the water. Take the temperature of the water in the flask. Measure the volume of water which has entered the flask. Find also the volume of the whole flask up to the lower surface of the stopper.

The method of calculation will perhaps be best understood by considering first an experiment in which the flask was opened under *ice-cold* water :—

Temperature of boiling water	100° C.
Temperature of cold water	0° C.
Volume of whole flask	360 c.c.
Volume of water entering flask	96 c.c.

Then,

360 c.c. of air contracted 96 c.c. in cooling through 100° C.

Therefore,

(360 − 96) c.c. or 264 c.c. at 0° would expand 96 c.c. when heated to 100° C.

Hence,	1 c.c.	,,	,,	$\frac{96}{264}$	,,	,,	100° C.
	1 c.c.	,,	,,	$\frac{96}{264} \times \frac{1}{100}$	,,	,,	1° C.
	1000 c.c.	,,	,,	$\frac{96}{264} \times \frac{1000}{100}$	,,	,,	1° C.
Or,	1000 c.c.	,,	,,	3·63 c.c.	,,	,,	1° C.

Secondly, consider an experiment in which the flask was opened under water at the *ordinary temperature* :—

Temperature of hot water	100° C.
Temperature of cold water . . .	16° C.
Volume of whole flask	360 c.c.
Volume of water entering flask	85 c.c.

Then,

300 c.c. of air in cooling from 100° to 16° (*i.e.* through 84°) contracted 85 c.c.
i.e. 300 c.c. in cooling through 84° contracted ... ... 85 c.c.

Hence,

360 c.c. in cooling through 1° would contract ... ... $\frac{85}{84}$ c.c.

360 c.c. ,, ,, 100° ,, ,, ... $\frac{85}{84} \times 100$ c.c.

Or,

360 c.c. would contract 101·2 c.c. in cooling through ... ... 100° C.

Therefore,

(360 − 101·2) c.c. or 258·8 c.c. at 0° would expand 101·2 c.c. when heated 100° C.

Hence, 1 c.c. ,, ,, ,, $\frac{101 \cdot 2}{258 \cdot 8}$ c.c. ,, 100° C.

1000 c.c. ,, ,, ,, $\frac{101 \cdot 2}{258 \cdot 8} \times \frac{1000}{100}$ c.c. ,, 1° C.

Or, 1000 c.c. ,, ,, ,, **3·91** c.c. ,, 1° C.

More accurate experiments than can be made with this simple apparatus show that the expansion of 1 litre of air at 0° when heated to 1° C. is 3·66 c.c. Dalton himself obtained the value 3·63.

Notice that we have assumed in these calculations that the expansion during 1 degree is the same wherever that degree is upon the scale of temperature. By repeating the experiment through different ranges of temperature, say 100° to 0°, 100° to 30°, 100° to 50°, we could find whether we are led in each case to the same result. Careful experiment proves this to be so, and therefore the assumption made is justified.

Suppose we have equal volumes of *different* gases, will they expand to the same extent as air, when heated through the same range of temperature?

EXPT. 2.[1] **To compare the Expansion of Carbonic Acid Gas with that of Air.**—Fit up the apparatus shown in Fig. 57. There are two similar tubes A and B within a wider one. A contains dry air, B an equal length of dry carbonic acid gas. They dip into and are partly fixed with dry mercury. The mercury is contained in a stout bottle C. The cork of this bottle is not air-tight. Surrounding the parallel tubes is a steam jacket D. The mercury levels stand side by side at E. Connect the

[1] Lecture Table Experiment.

tube F with a flask or tin can in which water is being vigorously boiled. The gases within the tube are heated by the steam. Observe that they both move down to the same level G. This shows that the expansion of the carbonic acid is equal to that of the air.

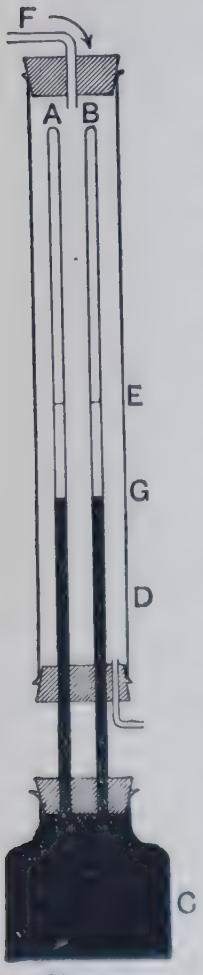

FIG. 57.

CHARLES'S LAW.—In the same way as in Experiment 2, the expansion of other gases may be compared with that of air, and the conclusion arrived at may be expressed in words used by Dalton in 1801 when he read his paper on this subject in Manchester:—

"The results of several experiments made upon hydrogenous, oxygenous, carbonic acid gas, and nitrous gas, which were all the kinds I tried, agreed with those of common air. . . . Upon the whole, therefore, I see no sufficient reason why we may not conclude that all elastic fluids under the same pressure expand equally by heat."

It was long forgotten that Dalton had first arrived at this important generalisation. Consequently it is after the Frenchman, who also independently arrived at it, that it is usually known under the name of *Charles's Law*. Viz.:—

Equal volumes of all gases under the same pressure expand equally for equal increments of temperature.

Accurate experiments show that 1 litre of all permanent gases at 0° will expand the same amount, viz., 3·66 c.c. for every rise of 1° C. Hence all gases expand $\frac{3{\cdot}66}{1000}$ or $\frac{1}{273}$ of their volume at 0° for every rise of 1° C. This number is called their **coefficient of expansion.**

For example,

273 c.c. at 0° become 273 + 1 c.c. at 1°
" 273 + 2 c.c. at 2°
" 273 + 3 c.c. at 3°
" 273 + 10 c.c. at 10°
" 273 + *t* c.c. at *t*°

We can, therefore, construct a Table such as the following, expressing the volumes of a gas at different temperatures which correspond to 273 volumes at 0° C. ; *e.g.*—

293	volumes at	20° C.
290	,,	17° C.
288	,,	15° C.
283	,,	10° C.
274	,,	1° C.
270	,,	− 3° C.
263	,,	− 10° C.

Absolute Zero.—Then, if on cooling a gas from 0° C. downwards, it contracts $\frac{1}{273}$ of its volume for every degree, we shall at last reach that temperature at which its volume entirely disappears. This will evidently be at − 273° C. *This temperature is called the absolute zero*, and it is frequently used as a convenient point from which to measure temperatures.

As a matter of fact, no one has yet succeeded in obtaining so low a temperature ; and most gases are liquefied above − 273° C. Still, *at ordinary temperatures a gas behaves as though its volume would vanish at* − 273° C.

Temperatures reckoned in Centigrade Degrees from Absolute Zero are called Absolute Temperatures (A), e.g.—

0° C.	corresponds to	273° A.
20° C.	,,	273 + 20 or 293° A.
− 15° C.	,,	273 − 15 or 258° A.
t° C.	,,	273 + t or T°.

If a gas has a volume 1 at 0° C. its volume at 10° C. will be $1\frac{10}{273}$

i.e., the volumes are as $1 : 1\frac{10}{273}$

or, $\frac{273}{273} : \frac{283}{273}$

or,

$273 : 283$

or, as their absolute temperatures.

In a similar manner it can be shown that the following restatement of *Charles's Law* is true :—**The volume of a gas is proportional to its absolute temperature, when the pressure is constant.**

Thus, if V_1 denote a volume of gas at the absolute temperature T_1,
and V_2 ,, ,, ,, ,, T_2,
then,

$$V_1 : V_2 = T_1 : T_2$$

or,

$$V_1T_2 = V_2T_1$$

EXAMPLES—(1) *Find the volume at 0° of 72 c.c. of a gas measured at 15° C.*

0° C. = 273° A. : 15° C. = 288° A.

If volume at 288° A. is 72 c.c.

at 1° A. it is $\frac{72}{288}$ c.c.

at 273° A. it is $72 \times \frac{273}{288}$ c.c.

i.e. at 0° C. it is $68\frac{1}{4}$ c.c.

Or, again, by reference to the Table on p. 133,
since, 288 vols. at 15° C. correspond to 273 vols at 0° C.,

therefore, 72 c.c. ,, ,, $72 \times \frac{273}{288}$ c.c. at 0° C.

i.e., 72 c.c. ,, ,, $68\frac{1}{4}$ c.c. ,,

(2) *At what temperature will 60 c.c. of gas measured at 27° C. expand to 80 c.c.?*

27° C. = (273 + 27)° A. or 300° A.

If the volume is 60 c.c. at 300° A.

it is 1 c.c. at $\frac{300^\circ}{60}$ A.

or 80 c.c. at $\frac{300}{60} \times 80^\circ$ A. or 400° A.

i.e., 80 c.c. at (400 − 273)° C. or 127° C.

See Examples III, p. 137.

Since a gas tends to expand when it is heated, it follows that if a gas is so heated that the volume cannot change the pressure must increase.

EXPT. 3.[1] **To show that when a Gas is heated at Constant Volume the Pressure is increased.**—Place a small piece of sheet india-rubber at the bottom of an 8 oz. round-bottomed flask. Introduce a test tube into it so that it rests on the rubber and is supported by the neck of the flask. (Take the temperature of the air in the flask.) Close the flask

[1] Lecture Table Experiment.

with a tightly-fitting rubber stopper, through which is passed a long tightly-fitting narrow glass tube (Fig. 58). Pour some mercury into the tube and see that it takes the same level within and without the tube. Measure the position of this level. Tighten the stopper. Surround the flask with boiling water and

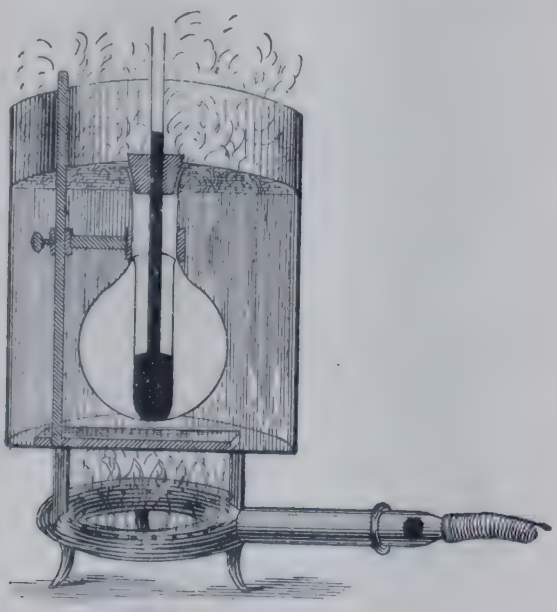

FIG. 58.

steam, as in Experiment 1. The mercury rises in the tube. Take the temperature of the water. Mark the level reached by the mercury in the tube. Measure the pressure shown by the mercury. For instance, if

Barometer stood at	750 mm.
Rise of mercury	240 mm.
Original temperature of the air	10° C.
Temperature of the boiling water... ...	100° C.

Then, for 90 rise of temperature, pressure rose 240 mm.

Hence, ,, 1° ,, ,, ,, would rise $\frac{240}{90}$ mm.

,, 10° ,, ,, ,, ,, $\frac{2400}{90}$ or $26\frac{2}{3}$ mm.

and at 0°, pressure would be $(750 - 26\frac{2}{3})$ mm.

$= 723\frac{1}{3}$ mm.

Therefore,

$$\text{coefficient of the increase of pressure for } 1^\circ = \frac{\frac{240}{90}}{723\frac{1}{3}} = \frac{1}{261} +$$

Thoroughly careful experiments show that when a gas is heated at constant volume the coefficient of the increase of pressure for 1° is $_2\frac{1}{7}\frac{1}{3}$. Observe that this is exactly the same as the coefficient of expansion when a gas is heated at constant pressure. Can you explain this apparently remarkable coincidence? Hence, *when a gas is heated at constant volume the pressure is proportional to the absolute temperature.*

Mathematical Expression for a Gas under different Conditions of Pressure, Volume and Temperature. Let us suppose that a gas in the state denoted by P_1, V_1, T_1, is converted into the state denoted by P_2, V_2, T_2, in *two* stages.

(1) While the temperature remains constant at T_1, change P_1 to P_2.

Let V_1 in consequence become V.

Then,

$$P_1V_1 = P_2V \qquad \text{(Boyle's Law)},$$

or

$$V = \frac{P_1V_1}{P_2}$$

(2) Then while the pressure remains constant at P_2, change T_1 to T_2.

Let V in consequence become V_2.

Then,

$$VT_2 = V_2T_1 \qquad \text{(Charles's Law)}.$$

But

$$V = \frac{P_1V_1}{P_2}$$

Hence,

$$\frac{P_1V_1}{P_2} \times T_2 = V_2T_1$$

or

$$\frac{\mathbf{P_1V_1}}{\mathbf{T_1}} = \frac{\mathbf{P_2V_2}}{\mathbf{T_2}} \quad \ldots\ldots\ldots \text{(1)}.$$

which is the expression required.

Observe that if we put $V_1 = V_2$ in equation (1), we have

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

or, when the volume remains constant, the pressure is proportional to the absolute temperature, as we have already shown by experiment to be the case.

EXAMPLE.—819 c.c. of air are measured at 13° C. and 741 mm. What would the volume be at 0° C. and 760 mm.?

Since 0° is a lower temperature than 13°, the volume will be less at 0°, and the volume corrected for temperature will be $819 \times \frac{273}{286}$.

Again, since 760 mm. is a greater pressure than 741 mm., the volume will be less at 760 mm. in the ratio $\frac{741}{760}$.

Hence,

$$\text{volume at } 0^\circ \text{ and } 760 \text{ mm.} = 819 \times \frac{273}{286} \times \frac{741}{760}$$
$$= 762{\cdot}3 \text{ c.c.}$$

N.B.—Gas volumes are generally corrected to 0° C. and 760 mm. of mercury.

N.T.P. means Normal Temperature and Pressure; *e.* 0 C. and 760 mm. of mercury.

EXAMPLES III

Charles's Law

1. Find from each of the following the temperature at which the volume of the air would vanish, if it went on contracting at the same rate (*see* Experiment 1)—

	(1)	(2)
Temp. of air when hot ...	100° C.	100° C.
Temp. of air when cold ...	0° C.	20° C.
Vol. of whole flask	414 c.c.	339 c.c.
Vol. of water entering... ...	111 c.c.	73 c.c.

2. Find the coefficient of expansion of air from each of the following—

	(1)	(2)
Temp. of air when hot ...	80° C.	99° C.
Temp. of air when cold ...	0° C.	16° C.
Vol. of whole flask	441 c.c.	338 c.c.
Vol. of water entering ...	100 c.c.	76 c.c.

3. A litre of air is measured at 13° C., what will be its volume at 0° C.? (Answer in c.c.)

4. What volume of air at 21° C. will fill one litre at 0° C.?

5. If a litre of air at 0° C. weighs 1·29 grams, what will be the weight of a litre at 28° C.?

6. A fire balloon contains 1,000 litres of air at 114° C. How much lighter is this air than 1,000 litres at 0° C.?

EXAMPLES IV

Charles's and Boyle's Laws

1. Find the volume at N.T.P. of—
   - (*a*) 100 c.c. of oxygen at 63° C. and 741 mm.
   - (*b*) 100 ,, 77° C. and 800 mm.
   - (*c*) 100 ,, − 52° C. and 676 mm.
2. A quantity of air measures 500 c.c. at 0° C. and 760 mm.
   - (*a*) Find the volume at 15° C. and 720 mm.
   - (*b*) ,, ,, 33° C. and 760 mm.
   - (*c*) ,, ,, 52° C. and 950 mm.
3. A quantity of nitrogen measures 20 c.c. at 17° C. and 740 mm.
   - (*a*) Find the temperature if the volume is 40 c.c. at 760 mm.
   - (*b*) ,, ,, ,, ,, 37 c.c. at 600 mm.
   - (*c*) ,, ,, ,, ,, 20 c.c. at 814 mm.
4. A quantity of carbonic acid measures 40 c.c. at 27° C. and 780 mm.
   - (*a*) Find the pressure if the volume is 40 c.c. at 0° C.
   - (*b*) ,, ,, ,, ,, 50 c.c. at 15° C.
   - (*c*) ,, ,, ,, ,, 30 c.c. at 6° C.

EXAMPLES V

The Densities of Gases

Assume—1 litre of hydrogen weighs ·09 gram at N.T.P.; also densities relative to hydrogen :—

Oxygen	16	Steam	9
Nitrogen	14	Carbonic acid ...	22

1. Find the weight of a litre of oxygen at N.T.P.
2. What volume will 1 gram of carbonic acid gas fill?
3. Find the volume occupied by 1 gram of steam at 0° C.
4. Find the weight of a litre of steam at 104° C.

5. Find the volume occupied by 169 grams of carbonic acid at 51° C. and 741 mm.

6. Find the weight of 19 c.c. oxygen at 13° C. and 770 mm.

7. To what temperature must oxygen (density 16) be heated in order to make it as light as hydrogen is at 0°?

8. At what temperature will air have half the density that it has at 23° C.?

CHAPTER XIV

A RESEARCH ON CHALK

We shall now return to Black's research[1] on chalk, and endeavour to follow those lines of inquiry which Black himself pursued. We may hope to arrive at important results by making in this way a thorough examination of one substance, particularly if we pay close attention to the quantities of the materials with which we are dealing.

Chalk is a substance which causes the liberation, as we have already found, of carbonic acid gas when hydrochloric acid is added to it, and yields lime when it is heated in a kiln. This is also true of calcspar and marble. Pending further evidence we shall use the term chalk as including all of these.

THE ANALYSIS OF CHALK

(*a*) The Action of Acids upon Chalk.—Expt. 1. **To find the Weight of Carbonic Acid expelled from 100 Grams of Chalk in contact with Hydrochloric Acid.**—We might attempt to collect and weigh the gas itself, but it is easier to find the *loss in weight* due to escaping gas. Choose a small lipped beaker and a small flask, small enough to stand side by side upon the balance pan.[2] Into the beaker

[1] A reprint of Black's admirable paper (1755) may be obtained in No. 1 of the "Alembic Club Reprints," W. F. Clay, Edinburgh, 1*s.* 6*d.* net. Advanced students and teachers of Chemistry are strongly recommended to obtain it. A few years later Lavoisier repeated and extended Black's experiments. He gave in 1773 a very clear account of them and of the history of the discovery of carbonic acid gas. A translation of this by Henry under the title of *Essays, Physical and Chemical*, is very readable and well deserves perusal.

[2] If there is not room for both on the pan, the flask may be slung by a thread from the hook at the end of the beam.

pour about 10 c.c. of strong hydrochloric acid. It fumes and may lose weight by fuming. Add about 30 c.c. of water ; the fuming stops. Place a short glass tube, which has been narrowed at one end, in the beaker. Weigh the half-full beaker and the empty flask together. Place about two grams of *dry* chalk[1] in the flask. Weigh the flask and beaker again (Fig. 59). Transfer carefully a little acid by means of the dropping tube from the beaker into the flask. There is an effervescence of carbonic acid gas. We wish to liberate all the carbonic acid ; so go on adding acid until effervescence ceases. By means of the tube suck out carbonic acid from the flask. Weigh the beaker and flask again together with the tube. There is a decrease in weight. Calculate the loss in weight there would be in the case of 100 grams of chalk. What causes of error are there possibly in this experiment?

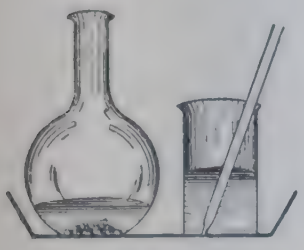

FIG. 59.

In successive experiments performed by pairs of boys the following numbers were obtained for the percentage of carbonic acid expelled from chalk and hydrochloric acid :—45, 42·4, 45·4, 43, 45, 43. *Mean* : = 43·9. Do your experiments lead to the same result?

Careful experiments show that whatever the quantity of the chalk that is taken, there is **a loss of about 44 per cent.** This shows that the quantity of carbonic acid evolved depends upon the quantity of chalk taken, and not upon the quantity of acid added, provided sufficient acid has been added. Is it therefore certain that the carbonic acid comes from the chalk and not from the acid (*vide* p. 9)? No.

(*b*) THE EFFECT OF HEAT UPON CHALK.—As we have already learnt, it has long been known that when chalk is exposed to a violent heat it is converted into a substance called **quicklime.** The alchemists and early chemists regarded this as a peculiar acrid and caustic earth, formed by the union of the chalk with a fiery substance.

In limestone and chalk districts you may have seen kilns in which quicklime is obtained by heating the stone with a little

[1] Use powdered calc-spar, or precipitated chalk, B.P., dried at 100°.

coal. The change can, however, be brought about by heat alone.

EXPT. 2. **To show that Chalk is changed by Heat Alone.**—(*a*) Damp a slip of red litmus paper. Let a little chalk dust fall upon it. The colour is unchanged.

(*b*) Heat strongly a little chalk dust on platinum foil with the Bunsen flame. Then drop it on the damp test-paper. The colour goes blue where the heated chalk has touched it. Then the chalk dust has, in part at any rate, been converted into quicklime.

It had apparently occurred to no one before the time of Black to really examine what changes were effected ; no one had thought of *weighing* the chalk before and after it was heated.

EXPT. 3. **To find the Weight of Quicklime obtained from 100 Grams of Chalk by Heat.**—Weigh a small porcelain crucible. Weigh into it accurately about one gram of pure chalk. Place it in a muffle furnace,[1] and heat strongly for half an hour. Remove the crucible from the furnace. Allow it to cool. Weigh. Reheat for ten minutes and re-weigh, and so on till there is no further change in weight. Here is the result of an actual experiment :—

Weight of crucible	13·42 gm.
Weight of crucible and chalk	14·63 ,,
Weight of chalk	1·21 ,,
Weight of crucible and lime, heated 30 mins.	14·10 ,,
,, ,, ,, ,, 30 + 10 mins.	14·10 ,,
,, ,, ,, ,, 30 + 10 + 10 mins....	14·10 ,,

∴ Weight of lime[2] = 14·10 − 13·42, or 0·68 gm.

$$= \frac{0{\cdot}68}{1{\cdot}21} \times 100, \text{ or } \mathbf{56{\cdot}2 \text{ per cent.}}$$

Pure chalk is always found to lose about 44 per cent. in weight, when ignited (*i.e.* strongly heated). Thus, instead of gaining anything from the fire, the chalk loses the *same* amount as when treated with an acid. It appears, therefore, from these experiments that the carbonic acid which is expelled from chalk

[1] Fletcher's Reverberatory Gas Furnace, p. H, is suitable for class work. A number of crucibles can be heated in it at the same time. If not available, a foot-blowpipe will give excellent results.

[2] The lime may be kept for Experiment 6.

and an acid is also driven off by heat, when chalk is converted into quicklime. It is therefore clear that in the former case the carbonic acid comes from the chalk.

If this is so, when acid is added to the quicklime there will be no effervescence.

EXPT. 4. **To try whether Quicklime contains Carbonic Acid.**—Drop hydrochloric acid on to a fragment of freshly burnt quicklime in a test-tube. The drops hiss as they touch the lime. Cover the lime with more acid. The liquid soon becomes hot and bubbles. Are the bubbles steam or gas? Cool the tube by plunging it into cold water. The bubbling stops. Try the experiment again using weak instead of strong acid. The mixture gets warm—not so hot as before, and there are no bubbles. The bubbles seem to have been steam, not gas; the mixture, then, did not effervesce, but boiled. Notice that the lime is gradually dissolved by the acid.

Quicklime, then, is chalk which has been deprived of its carbonic acid.

CONCLUSION FROM THE ANALYSIS OF CHALK—

Now, 100 grams chalk lose on solution in acid 44 grams in weight, and at the same time carbonic acid is evolved.

100 grams chalk lose on ignition 44 grams in weight.

100 grams chalk yield on ignition 56 grams quicklime, which contains no carbonic acid.

But $44 + 56 = 100$.

Therefore, there seems good reason to believe that

chalk = lime + carbonic acid gas,[1]

nothing more and nothing less.

The Properties of Quicklime and of Slaked Lime.—Preparation of Pure Quicklime.—We shall find presently that quicklime deteriorates when exposed to the air. On this account quicklime, when not fresh, should be reburnt before experimenting with it.

Put some lumps of quicklime into the middle of a hot coal fire. Leave them there red-hot for several hours. Then remove them.

[1] The sign = means "yields" or "gives." It has not the same meaning as in algebra and chemical equations (unlike algebraic equations) read only from left to right.

To test the purity of the lime add water and hydrochloric acid. There should be no effervescence.

Place the rest in tin boxes with closely fitting lids, for future use. Seal with paraffin. Even then the quicklime should be reburnt before any exact experiment. The lightest pieces are generally the best. Heavy and stony pieces should be rejected.

EXPT. 5. **To Slake Quicklime.**—It has long been known that water has an action upon quicklime. Place a lump of quicklime upon a plate. Pour a little water over it. Observe that it presently crumbles up: much heat is evolved, for steam is produced. The quicklime is said to be *slaked*, or converted into slaked lime. What has occurred? Suppose we ask the balance.

EXPT. 6. **Is Slaked Lime Quicklime + Water?**—Take a small crucible. Weigh it. Introduce about 2 grams of good fresh quicklime. Weigh. Carefully moisten it with about 1·5 c.c. of distilled water. Place the crucible in a steam-jacketed oven, or in an oven kept at 100° C., for an hour, to evaporate excess of water. Weigh. Reheat at 100°, and reweigh, till no further decrease in weight occurs. In an actual experiment—

Weight of crucible	13·33 gm.
Weight of crucible and quicklime	14·01 ,,
Weight of quicklime	0·68 ,,
Weight of crucible and slaked lime	14·23 ,,
,, ,, ,, ,, (re-dried) ...	14·23 ,,

Gain in weight of slaked lime = 0·22 gm.

,, ,, ,, $= \frac{0{\cdot}22}{0{\cdot}68} \times 100$ or **32·3** per cent.

Then, when quicklime is slaked with water, it *gains* in weight.

When good quicklime is carefully slaked, there is always found a gain in weight of about 32 per cent. This must be due to water uniting with the quicklime.

Hence,

Quicklime + water = slaked lime.

We found some time ago that slaked lime is slightly soluble in water. But we have not obtained as yet an answer to the following question.

EXPT. 7. **Can Slaked Lime be entirely dissolved in a Large Quantity of Water?**—Why does not the white mud of slaked lime entirely dissolve in water? Is the mud which settles something different from the lime which dissolves—*i.e.* is the lime impure? Does it need heat, or shaking, or longer time? Try the experiment again, giving the lime every chance. Choose a good piece of lime—one which feels light—or one of the lime cylinders used for the lime-light. Burn it in a coal fire, and when still fresh take *not more* than 0·5 gram. Have at least 500 c.c. of freshly boiled *distilled* water ready in a flask. Slake the lime with a few c.c. of boiling hot distilled water. Throw the mixture at once into the flask of hot water. Shake.

Black tried this experiment, and found that the whole of the lime dissolved with the exception of a very small amount of sediment. This, as Black observed, was probably due to the accidental admixture of a little foreign matter.

Is Quicklime an Element?—Is quicklime an elementary substance, *i.e.*, an element? So far as our present study of the subject goes, and so far as Black knew, it is an element, for we do not know how to break it up into different substances. But observe that this does not prove that it is really an element. Indeed, quicklime resembles magnesia (*i.e.* oxide of magnesium) in appearance and properties. Both are infusible, white, earth-like substances; and magnesia can be obtained from another white substance, "magnesia alba," in just the same way as quicklime from chalk. Up to the time of Black, these substances were frequently confused with one another. We know now that magnesia is really an oxide of the metal magnesium. It is possible, then, that further study or research may prove quicklime to be an *oxide* of a metal.

THE SYNTHESIS OR PREPARATION OF CHALK FROM LIME.—We have shown by analysis that chalk is a compound of lime and carbonic acid gas. Can we not, then, perhaps produce chalk by uniting carbonic acid with lime?

EXPT. 8. **To find whether Chalk can be formed from Lime and Carbonic Acid.** Powder some quicklime. To part of the powder add water, and then hydrochloric acid. If it does not effervesce, it is free from chalk, and will do for this experiment.

(1) Weigh a small dry flask. Introduce about 1 gram of quicklime. Weigh. Lead into the centre of the flask a current of dry carbonic acid gas, drying the gas by bubbling it through some strong sulphuric acid in another flask. After 10 minutes remove the flask, suck out the gas, and weigh. Is there any change in weight? Probably little or none. Add acid: there is no effervescence. Hence, no chalk has been formed.

(2) Place about 1 gram of the quicklime in a test-tube. Half fill the tube with water. Bubble carbonic acid gas up through the milky liquid for a few minutes. Pour the whole of the contents of the tube upon a filter. Wash two or three times with water. Drop 2 or 3 drops of hydrochloric acid on to the white substance on the paper. It effervesces violently. Then chalk has probably been formed.

*(3) How could it be proved beyond any doubt that it is chalk, and not anything else containing carbonic acid, which has been formed? Would it not be well to find whether the product contains the same percentage of carbonic acid that we know to be present in chalk? Proceed thus. Grind up a little quicklime with water, and pour off the milky liquid into a flask, and pass in carbonic acid gas for at least half an hour. Boil for a few minutes. Filter. Dry the filter-paper with the precipitate in a steam oven at 100°. Heat some of the dry substance in a crucible as in Experiment 3, and find the percentage loss of weight. It is nearly 44 per cent.

The above experiments show that water must be present if chalk is to result from lime and carbonic acid gas. Does chalk then contain water? No, the fact that the loss in weight of chalk when acted on by aqueous acid is the *same* as that incurred when strongly heated makes it certain that chalk does not contain water. We must, however, suppose that water is in some way necessary for the union of lime and carbonic acid gas, though the water does not itself enter into combination.

We have, then, succeeded in forming chalk by putting together lime and carbonic acid gas. Expressed briefly, in addition to the previous conclusion that

Chalk = lime + carbonic acid,

we have now shown that

Carbonic acid gas + lime = chalk

This is another instance of what is called a *synthesis*, or placing together. It affords most important additional evidence of the nature of chalk, and confirms the conclusion we had arrived at from the *analysis* of chalk.

Carbonates.—Chalk, as a substance produced from the alkali lime and from the acid carbonic acid, may clearly be regarded as a *salt* (cf. Vol. I. p. 167). And just as salts formed from nitric acid are called nitrates, so salts derived from carbonic acid are called **carbonates**. Chalk may therefore be called "*carbonate of lime*." Any other substance which evolves carbonic acid gas on addition of an acid is called a carbonate; soda crystals, for example, may be termed a carbonate of soda.

The last experiment renders it extremely probable that it is chalk which is produced when carbonic acid is passed through clear lime water. (*See* p. 7, Problem 1).

EXPT. 9. **The Lime Water Test for Carbonic Acid.**—Pour a little lime water into a jar containing carbonic acid gas. Shake. It becomes milky. A slight precipitate forms. Show that it is chalk.

Clear lime water, as we found previously, may be used as a delicate test for carbonic acid. We now understand that the lime water is rendered milky because *chalk* is at once formed, which is insoluble in water. One drop of lime water is often enough. For example:—

EXPT. 10. **To show that Carbonic Acid is given off when Chalk is heated.**—Place a few grams of chalk in a hard glass test-tube. Heat the tube in the foot blow-pipe flame. Dip a glass rod into lime water. One drop of lime water hangs at the end of the rod. Carefully lower the drop into the test-tube, so that it does not touch the sides. Withdraw the rod. Examine the drop against a dark background. It has become cloudy.

We have already noticed that when lime water is exposed to the open air a pellide or crust forms upon its surface (*see* p. 8, 4). As Black said in 1755:—

"If this fluid be exposed to the open air, the particles of quicklime which are nearest the surface, gradually attract the particles of Fixed Air which float in the atmosphere. But at the same time that a particle of lime is thus saturated with air, it is also restored to its native state of mildness and of insolubility,

and as the whole of this change must happen at the surface, the whole of the lime is successively collected there, under its original form of an insipid calcareous earth, called the cream or crust of lime water."

Carbonic acid gas can always be found by this test in the atmosphere, though usually only in very small quantities (cf. p. 8).

Can chalk be prepared from lime in any other way? Now, soda crystals (mild alkali) contains carbonic acid (*see* p. 9, Problem 1). What then may the effect be of adding quicklime to a solution of soda crystals? May not the lime recover carbonic acid from the soda? This is a question Black asked himself. We may answer it by experiment, as he did.

EXPT. 11. **To find whether Chalk can be re-formed from Lime by means of Soda Crystals.**—(1) Weigh out about 7 grams of soda crystals. Dissolve the soda in about 60 c.c. water. Filter the solution into a flask. Powder finely about 1 gram of pure quicklime. Throw the quicklime into the solution of soda. Warm and shake well. Pour the product on to a filter-paper. Wash the "precipitate" on the paper repeatedly with hot water, so as to free it completely from the soda. Pour a few drops of acid on to the paper. There is a violent effervescence. The precipitate contains, therefore, ordinary chalk. Then the carbonic acid which the lime has gained must have been furnished by the soda crystals.[1]

(2) Dissolve a few soda crystals in a little water, and add the solution to some clear lime water. The clear solution becomes milky. Why? Because a precipitate is formed, which if collected in sufficient quantity may be shown to be chalk. Why is the precipitate so small in quantity?

*EXPT. 12. **To find the Weight of Chalk obtained from 56 Grams of Quicklime dissolved in Acid.**—Weigh out accurately into a flask 1·5 to 2 grams of pure quicklime. (From what quantity of chalk could it be obtained?) Add water and hydrochloric acid. If there is no effervescence, it proves that the lime is free from gas. If there is any effervescence, then any gas contained as an impurity in the lime is

[1] Advanced students may determine the quantity of chalk produced by this means from a weighed quantity of quicklime. The precipitate should be dried in an oven at 100°.

got rid of. The lime will dissolve on warming. If it does not all disappear, add a little more acid. The solution now contains excess of acid. Add to it a clear solution of soda crystals. There is an effervescence. (Why?) Then a precipitate forms. But on shaking, this precipitate disappears with effervescence. (Why?) After adding still more soda the effervescence stops (why?), and the precipitate remains undissolved. Go on adding more soda until no more precipitate forms. To know when there is enough soda, filter a little of the liquor containing the precipitate through a counterpoised filter-paper. Add soda to the filtrate. If there is no precipitate, then the soda is already in excess.

The precipitate must now be obtained clean from the soda, hydrochloric acid, and other substances which may be present. Heat the liquid until it *boils*. The precipitate clots and settles. Pour the liquid on to the filter used before. Wash the precipitate again with hot water, and pour the washings on to the filter. Pour the precipitate on to the filter. Go on washing it with hot water, until it is quite free from soda or other dissolved substances. Put one drop of the filtrate as it drips from the funnel on to platinum foil and evaporate. If it leaves a white mark, the precipitate is not yet clean, and must be washed with more water.

Dry completely the filter-paper with the chalk in the funnel at 100° in an oven. Carefully place the paper with the chalk on the pan of the balance, place the counterpoise of the filter on the other pan, and weigh again. Find the weight of the chalk alone. Hence, calculate the weight of chalk obtained from 56 grams of quicklime. Compare your results with those obtained in Experiment 3.

EXPT. 13. **An Attempt to recover Lime from its Solution in Acid.**—Lime dissolves in acid. Chalk when dissolved in acid loses carbonic acid gas. Can we recover the lime by evaporating the acid?

Leave about 2 grams chalk soaking in hydrochloric acid so that most of the chalk (but not quite all) is dissolved. Filter the liquid. Evaporate in a porcelain basin to dryness.

Is the solid white substance lime? Try whether a portion forms lime water with water. No! It dissolves much more

readily. Is the solution alkaline? No! Notice that the solid soon becomes sticky and moist if left exposed to the air. Evidently it has the power of attracting moisture from the air.

This new substance is called "**calcium chloride.**" Its most remarkable property is the readiness with which it picks up moisture. Hence, it is a useful substance for drying gases which may be passed over it, and indeed it has already been employed for this purpose (cf. p. 74, Expt. 11; p. 91, Expt. 4; p. 104, Expt. 4).

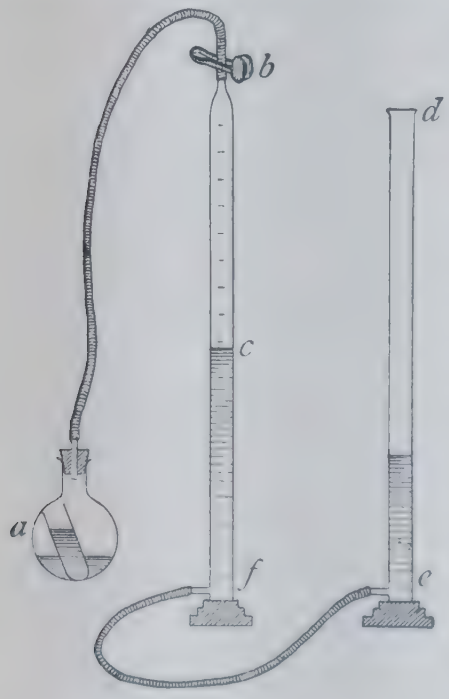

FIG. 60.

Thus, lime, which has been dissolved in hydrochloric acid, cannot be recovered by evaporation of the solution. On the contrary, we obtain a substance which is very different in its properties from both the alkali lime and hydrochloric acid. It will be shown later that it belongs to the class of substances called *salts*.

The Volume of Gas expelled from Chalk. — We have found that a definite quantity of chalk loses, on ignition or on solution in acids, a fixed *weight* of carbonic acid gas. Hence, a fixed *volume* of the gas should be obtained from a definite quantity of chalk.

EXPT. 14.[1] **To find the Volume in c.c. of Carbonic Acid Gas expelled from 1 Gram Chalk by Hydrochloric Acid.** — Weigh out carefully about 0·25 gram of chalk upon a piece of paper; fold this loosely, and introduce it into a 2 oz. flask, *a*. Add about 5 c.c. of water. Carefully introduce a small glass tube containing about 5 c.c. of dilute hydrochloric acid suspended by a thread; close the flask with a tightly fitting india-rubber stopper, through which is passed a glass tube connected with two feet of thin rubber tubing. Then

[1] Lecture Table Experiment.

attach the free end (Fig. 60) of the tubing to a pair of Hempel's burettes filled with water.[1] Place the burette *de* at a lower level than the other, and then open the clamp *b*. Tilt the flask so that the acid is brought into contact with the chalk. Carbonic acid is immediately evolved and forces the water down the tube *bc*. If the flask is gently shaken, the chalk will entirely dissolve.

Finally raise the burette *de* until the water stands at the same level in both columns. Read off the volume of air or gas in *bc*, and calculate the volume of carbonic acid which 100 grams of chalk would evolve. In an experiment :—

0·25 gm. chalk evolved 55 c.c. carbonic acid.

Hence,

1 ,, ,, 220 c.c. ,,

The experiment should be repeated several times. It is found that at 15° C. about 220 c.c. of carbonic acid are always obtained from 1 gram of chalk.

Problem

Fit up the apparatus shown in Fig. 32. Find how to use it so as to measure the volume of carbonic acid gas expelled from chalk when dissolved in dilute hydrochloric acid. Take about 1 gram of chalk. Is there any source of error?

CHALK AND THE HARDNESS OF WATER.—We observed some time ago that a "soft" water is rendered "hard" if shaken with chalk, and that spring waters in chalk districts are usually very hard (cf. Vol. I. p. 125). Let us examine this further, and attempt to obtain a hard water by passing carbonic acid gas into lime water, when, as we have found, chalk will be formed (see Experiment 9).

Expt. 15. **To pass Carbonic Acid Gas into Lime Water.**—Place about 200 c.c. of clear lime water in a flask, and bubble carbonic acid gas into it; the solution of course becomes milky, through the formation of carbonate of lime, *i.e.* chalk. But as the current of gas is continued, the white liquid unexpectedly

[1] This consists of a graduated glass tube, *bf*, connected at its lower end by a rubber tube with another glass tube, *de* (Fig. 60), open at both ends. The tube *bf* is narrowed at its upper end, and to it can be connected a piece of rubber tubing, on which is placed a clip, *b*. Open the clip *b*, and pour water into *d* until it flows out from *b*; then close the clip, and pour a little water out of the tube *de*.

becomes perfectly clear again. How can this be? Is the chalk destroyed, or is it simply dissolved? Let us examine the clear liquid which has been produced.

EXPT. 16. **To try the Effect of heating the Clear Solution.**—Heat some of the solution in a clean test-tube, and watch the solution closely. Hundreds of minute bubbles of gas rise to the surface, and the solution gradually becomes milky once more. Is the milkiness due to the appearance of lime or of carbonate of lime, *i.e.* chalk? Try: it is chalk.

Why has this chalk appeared in what was a clear solution? Scarcely owing to concentration of the solution, for very little water has been evaporated. Or is chalk less soluble in hot than in cold water? If so, the milky solution when cooled should become clear again. Try: the solution remains milky.

Recall the chief facts observed: firstly, when carbonic acid gas was bubbled through lime water, a milkiness at once appeared, due to the formation of chalk, which soon disappeared as the current of gas continued; and secondly, when the resulting clear solution was boiled, little bubbles of escaping gas were observed, and chalk reappeared. The disappearance of the chalk must clearly be connected with the solution of carbonic acid gas in the liquid. In other words, chalk will only dissolve in water if the water has also dissolved carbonic acid gas.

EXPT. 17. **To try the "Hardness" of the Solution.**—If we have correctly accounted for the facts observed, the clear solution prepared in Experiment 15 should be very "hard," and it should be greatly softened through being boiled.

(1) Find the number of c.c. of soap solution necessary to yield a permanent lather with 50 c.c. of the clear solution prepared in Experiment 15. Proceed as in Vol. I. p. 124, Expt. 8. The solution is found to be extremely hard.

Natural waters always contain carbonic acid gas in solution, derived from the atmosphere or from the soil, where it is constantly being formed from decaying vegetable matter. Consequently it is found that limestone strata are gradually eaten away into holes and caves, and that spring and river waters in chalk and limestone districts are very hard.

(2) Boil about 70 c.c. of the clear solution prepared in Experiment 15, for about five minutes, so as to throw out of solution or "precipitate" the dissolved carbonate of lime. Then filter, and

test the hardness of 50 c.c. of the clear filtrate. Do you find the water much softer?

In an experiment—

50 c.c. clear solution, unboiled, took 41 c.c. soap solution.
" " " boiled, " 2 c.c. " "

EXPT. 18. **To try the Effect of adding Lime Water to the Clear Solution.**—What visible effect is likely? Try adding some clear lime water to some of the solution of carbonate of lime prepared in Experiment 15. A white precipitate is at once formed. What has happened? The lime in the lime water has united with the dissolved carbonic acid, forming fresh carbonate of lime, and this has appeared as a white precipitate along with the carbonate of lime previously in solution, which can no longer remain dissolved when the carbonic acid is removed. If the proper amount of lime water is added, and the precipitated carbonate of lime removed, the hardness is, of course, greatly diminished. Indeed, on these facts is based **Clarke's Process for softening hard waters.**

Besides carbonate of lime, there is another substance—carbonate of magnesia—which is also dissolved by water containing dissolved carbonic acid gas; it also is precipitated when the solution is boiled or when lime is added to it. Hard waters which can be softened in either of these ways are said to possess **temporary hardness.**

On the other hand, there are some substances, particularly gypsum and Epsom salts (sulphates of lime and magnesia), also rendering waters hard, which are not precipitated either by boiling or by adding lime. The hardness of such waters is called **permanent hardness.** Such hardness can, however, be removed by distillation, or by the use of soda crystals (see Vol. I. p. 126).

When a water which contains carbonate of lime in solution is exposed to the air, much of its dissolved gas gradually escapes, and the dissolved carbonate of lime is therefore thrown out of solution. It is through this action that objects are *petrified* at Knaresborough, and *stalactites* are formed from the roofs of caverns, as in Clapham Cove, and as may often be observed from the vaulted roofs of bridges.

THE COMPOSITION OF WASHING SODA (SODA CRYSTALS).—We have already found that washing soda (soda crystals) contain water of crystallisation and carbonic acid.

We have also found that chalk is precipitated when lime is mixed with a solution of washing soda, *i.e.* of "mild alkali," as Black called it (see Experiment 11). When chalk is produced in this way, carbonic acid is transferred from the washing soda to the lime. What is left of the washing soda? It was stated in Vol. I. p. 159, that the substance called "caustic soda" could be prepared by these means. But until we have ourselves prepared it, suppose we call this other constituent of the washing soda X. Then :—

Washing soda = X + carbonic acid.

Then we can also express the change which occurs when lime is mixed with a solution of washing soda thus :—

Lime + washing soda = chalk + X.

i.e.

Lime + (X + carbonic acid) = (lime + carbonic acid) + X.

This is a chemical equation. What we require to do is to obtain the substance X.

EXPT. 19. **To obtain the Other Constituent of Washing Soda.**—Weigh out about 10 grams of quicklime. Slake it with boiling water. Then stir it up with more water so as to form a milky liquid of about 50 c.c. volume. Dissolve about 10 grams washing soda in 50 c.c. hot water. Add the solution to the milky liquid, stirring at the same time. Boil for ten minutes, and replace from time to time the water which evaporates. The white mud may be lime changed to chalk, or it may be partly lime and partly chalk. We must keep in mind that besides the visible mud there may be other substances present in solution. Washing soda may be in solution ; lime may be in solution. The substance X may be either in the solution or in the mud. Is it likely that both washing soda and lime are in solution together? No, since their solutions yield a precipitate when mixed (see Experiment 11).

Let us find whether the solution contains any washing soda. Remove the flame. The mud begins to settle. Filter a few

drops of the liquid. Add a drop of hydrochloric acid to the filtrate. It does not effervesce. Then there is no washing soda in the solution. (If there is any effervescence the liquid must be boiled longer.) The precipitate is so bulky, that it will hardly separate from the liquid. Pour both into a tall jar. Add an equal volume of water. This dilutes the solution. But on settling the precipitate occupies relatively less room, and allows the greater part of the solution or lye to be withdrawn as a clear liquid.

EXPT. 20. **The Lye is Caustic.**—Place some of the lye in a smooth-surfaced porcelain crucible or basin. Evaporate by heating the crucible on a sand-bath, surrounding it well with the sand. Cool. It solidifies as a white mass. Re-dissolve in water and wash out the basin. Observe that its inner surface is corroded. As Black said, "The acrimony of the lye is so great, that having once evaporated a part in a bowl of English earthen or stone ware . . . . it corroded and dissolved a part of the bowl, and left the inside of it pitted with small holes." On this account it is customary to evaporate the lye, when desired, in a silver or iron[1] dish, which are found to remain unattacked.

If a little of the lye is placed upon the skin, it will quickly raise a blister. On account of these properties it was called by Black and others "**caustic alkali,**" in distinction from "mild alkali" (washing soda). It is now called "**caustic soda.**"

Have we proved that X is caustic soda alone?

When mild alkali (washing soda) is added in sufficient amount to lime dissolved in hydrochloric acid, chalk is precipitated, as Experiment 12 showed. What will be the result of using caustic alkali in place of the mild alkali? Chalk cannot be formed. Why not? But the caustic alkali will neutralise the acid in which the lime is dissolved. Then perhaps the lime will be precipitated? Try.

EXPT. 21. **To find whether Lime is precipitated when Caustic Alkali is added to Lime dissolved in Hydrochloric Acid.**—Place about 1 gram of pure quicklime in a flask. Add water and hydrochloric acid. Warm till the lime is dissolved. Add a clear solution of caustic soda. A white precipitate is formed. Add the caustic soda until no further

[1] Platinum or aluminium dishes must not be used.

precipitate is formed. Heat until the liquid boils. Wash by decantation, and then pour the contents on to a filter-paper, as in Experiment 12. Wash thoroughly with hot water, to free it from caustic soda. Dry it at 100° in an oven. Examine the product. Lavoisier performed this experiment, and wrote :—

"This earth was a true lime ; it dissolved in water, in the same proportion as lime. The lime water prepared with it threw up a cream of chalk on its surface ; it made scarcely any effervescence with acids ; it communicated causticity to alkalies ; in short, no difference could be perceived between this and real lime prepared by calcination."

Summary.—The results of our study show that quicklime and caustic soda are less composite substances than chalk and washing soda. Yet the very reverse was supposed to be the case before the work of Black, which we have been closely following. For the acrimony of the caustic substances was supposed to be due to the presence of a fiery substance, absent from the mild substances. Black sums up the matter thus : "If quicklime be mixed with a dissolved mild alkali, it shows an attraction for Fixed Air superior to that of the alkali. It robs this salt of the air, and thereby becomes mild itself ; while the alkali is thereby rendered more corrosive" ; or, as we should say,

Lime + Washing Soda = Chalk + Caustic Soda.

Problems

1. Are the substances X, Y, Z, chalk or lime, or both or neither ?
2. Invent a substitute for the Hempel burette, using the ordinary apparatus of the laboratory.
3. Find whether there is any loss in weight when the given sample of quicklime is dissolved in hydrochloric acid. If there is a loss in weight, what inference do you make ?
4. Find whether lime can be made by heating chalk in an ordinary coal fire.
5. What weight of slaked lime can be dissolved by a litre of water ?
6. Find the percentage loss of weight when the given sample of dry slaked lime is strongly heated.

7. Find whether the given sample of dry slaked lime contains any carbonate of lime, and if so, how much?

8. Find whether the given sample of "quicklime" contains (*a*) any slaked lime, (*b*) any quicklime, (*c*) any carbonate of lime.

9. Find the effect of adding lime water to water saturated with carbonic acid gas and reddened by litmus.

10. Find whether carbonic acid gas can be driven out of soda crystals by heat.

11. Find what has occurred when clear crystals of washing soda have turned white through exposure in the air.

12. Find whether you can recover caustic soda from its solution in hydrochloric acid.

13. Find whether the given sample of caustic soda contains (*a*) any carbonate of soda, (*b*) any lime.

14. You are provided with some caustic alkali (caustic soda). Endeavour to prepare some mild alkali (washing soda) from it.

15. Find whether a solution of bicarbonate of soda is changed by being boiled. Is any carbonic acid gas given off?

16. Find in the case of (*a*) crystals of washing soda, (*b*) carbonate of magnesia, (*c*) carbonate of copper :—

(i) The percentage loss in weight when treated with hydrochloric acid.

(ii) The percentage loss in weight when strongly heated.

(iii) The effect of hydrochloric acid upon the substances obtained by ignition as in (ii).

(iv) The volume of carbonic acid expelled from 1 gram on treatment with hydrochloric acid.

17. Find the percentage weight of carbonic acid which can be obtained from Dover chalk, mountain limestone, coral, oyster-shells, egg-shells.

18. Prepare a sample of carbonate of magnesia from magnesia.[1]

19. Prepare a sample of carbonate of copper from oxide of copper.

20. Dissolve some iron nails in hydrochloric acid. Filter from any undissolved matter. Find the effect of adding to the solution (*a*) a solution of washing soda, (*b*) a solution of caustic soda. Collect, dry, and examine the products.

[1] *i.e.* oxide of magnesium.

CHAPTER XV

FUELS AND FOOD-STUFFS

It may well be asked why Fuels and Food-stuffs should be classed together. What connection is there between coal and bread, or between a candle and beef or beans? Our answer will be obtained by studying the effect of heat upon them.

Expt. 1. **To find the Effect of Heat upon Bread.**—(1) Place a small piece of dry bread in a dry test-tube, and heat it cautiously. A dew as of water soon appears on the upper and cooler part of the tube, and the bread changes colour, becoming yellow or brown, and at the same time heavy fumes are given off, which discolour the side of the tube, and burn if a light is applied to the mouth of the tube. Heat the bottom of the tube strongly. At last the fumes almost cease, and a black charred mass is left. Examine it. It is brittle and porous, and will mark paper. Is it carbon? If so, it should glow when strongly heated in the open air.

(2) Heat a small piece of bread in an iron spoon.[1] The result is at first the same, and a charred residue is left, but if the spoon is strongly heated this begins to glow and burn away, leaving at last only a little grey ash.

But, in order to prove certainly that the black residue is carbon, it will be worth while to find whether carbonic acid is formed when it is burnt.

Expt. 2. **To find whether the Charred Residue forms Carbonic Acid when burnt.**—One method of trying this would be to heat some of the charred residue within a tube in a

[1] A small square piece of platinum foil held by a crucible tongs may be used with advantage, or even an inverted crucible lid.

current of oxygen, passing the issuing gases into lime water. But it will be simpler to heat the substance with some solid substance, such as copper oxide, which we know will, in certain circumstances, yield up the oxygen it contains (see p. 103, Experiment 3). Proceed as follows :—

Char some bread by toasting it for a short time in an iron spoon. Grind a little of the charred mass with five or six times its weight of copper oxide, previously dried by heating it in a porcelain basin. Place the mixture in a small hard glass test-tube, fitted with a cork and delivery tube, bent once at right angles (Fig. 61). Clamp the tube horizontally, and let the delivery tube dip into a test-tube containing some clear lime water. Heat the hard glass test-tube strongly with a Bunsen burner. Presently, as the bubbles of gas pass through the lime water, it becomes milky. (Take care that the lime water does not run back into the hot tube.) Finally, empty the contents of the tube into a porcelain basin : red particles can be seen : some of the black oxide of copper has been *reduced* to metallic copper, while the charred mass has been *oxidised*, with formation of carbonic acid gas.

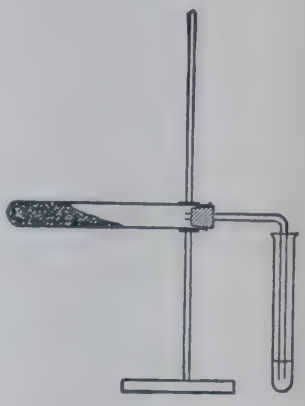

FIG. 61.

This experiment proves completely that **bread contains carbon.**

PROBLEMS

Find whether the following contain carbon :—

(*a*) *Foods*—

Starch[1]	Lean meat	Oatmeal	Rice
Sugar	Mutton fat	Peas	Hay
Biscuit	Dried egg albumen	Beans	

(*b*) *Fuels*—

Sawdust	Coal	Paper	Peat

It was discovered so long ago as the seventeenth century by Van Helmont that "vegetables by the action of fire" liberate the

[1] Starch is the chief constituent of the potato, wheat-flour, and rice.

gas which we now call carbonic acid gas, but he called it simply "*gas*," and knew nothing about its composition.

The fact that all the fuels in the above list, and they are all *solid* fuels, are proved to contain carbon, suggests that *liquid* fuels may also contain it. We may conveniently test liquid fuels in the following way.

EXPT. 3. **To find whether Alcohol contains Carbon.** —Take a clean *dry* gas jar, and a clean deflagrating spoon. Place a few drops of alcohol in the cup, light it, and introduce into the jar. A dew is deposited on the sides, and soon the flame dies down. Remove the spoon, and pour in some clear lime water and shake. It at once becomes milky, and it is therefore proved that alcohol also contains carbon.

PROBLEMS

Find whether the following contain carbon :—

Benzene	Ether	A wax candle	Coal gas
Olive oil	Naphtha	A tallow candle	

It may be remarked that Dr. Priestley, about 1772, first discovered that the gas we call carbonic acid gas resulted when candles, spirits of wine (*i.e.* alcohol), and ether were burned in a confined volume of air. He, in the language of the day, called the gas Fixed Air, but he could not see the importance of his discovery.

Every one of the substances we have examined is either a food-stuff or a fuel. The materials of some are formed by the processes of animal life, those of others by the action of vegetable life, and on these grounds they are often termed **organic**[1] substances.

But we have found that all these diverse substances have this at least in common, that they **all contain carbon**—they are all built up on a basis of carbon. And it is mainly because foods contain carbon that they have been found to be combustible ; they are therefore fuels, though they are too valuable to be used as fuels. It is evident, moreover, that carbon is one of the most common and most important of all elements, since

[1] Greek *organon*, an instrument, or that which works. "An organ is a part of a living body—the structure which fits it for the performance of some special action which is called its function" (Huxley).

some kind of fuel or of food is being slowly formed almost everywhere on the surface of our earth through the action of the vital processes of animals and vegetables.

When dry bread was heated in a test-tube, a liquid like water was deposited on the upper part of the tube, and when alcohol was burnt in a gas jar a dew was deposited. The same observation may be made with other fuels and food-stuffs. Is the liquid water? If so, these substances must contain *hydrogen*.

Expt. 4. **To find whether Sugar contains Hydrogen.**—Place some dry sugar[1] in a dry test-tube, and fit it with a delivery tube as in Experiment 2, dipping into a dry test-tube. Heat the sugar very gently, until a few drops of clear liquid are obtained. Find its boiling-point and freezing-point, proceeding as in Experiment 5, p. 92. It is proved to be *water*. **Sugar, therefore, contains hydrogen**, as well as carbon.

A sufficient quantity of the liquid formed when alcohol is burnt can be easily obtained by burning the alcohol in a small "spirit" lamp, fitted with an asbestos[2] wick, under a cold surface, as in Experiment 4, p. 91, and it can be proved to be *water*.

Problems

Find whether the following contain hydrogen:—

Starch[3]	Egg albumen	Paraffin oil
Flour	Coal gas	A candle

Many Organic Substances also contain Oxygen.—All food-stuffs and some fuels, when heated in absence of air or of any source of oxygen, liberate water; sugar and starch, for example, do. It follows that they contain **oxygen** as well as **carbon and hydrogen.**

Accurate experiments show that when sugar or starch are completely burnt, only carbonic acid gas and water are produced, and it follows that they contain no other elements besides carbon, and the elements of water, viz., hydrogen and oxygen. Moreover, if sugar or starch are completely burnt, and arrangements are

[1] Fine sugar may be dried in an air oven at 100° C.

[2] To avoid the objection that the hydrogen might be derived from the combustible cotton wick.

[3] It may be necessary to mix some of these with dry copper oxide.

made to catch and weigh[1] the carbonic acid and water which are produced, it is possible to calculate the percentage of carbon, hydrogen, and oxygen in the sugar or starch, and it is found that the hydrogen and oxygen are contained in the ratio in which they are present in water. On this account sugar and starch are called **carbohydrates** (*carbon*, and Gk., *hudor*, water). Other representatives of this class are the *gums* and *cellulose*,[2] which is the chief constituent of cotton and paper.

A Further Comparison between Fuels and Foodstuffs.—If the common presence of carbon is the only connection between fuels and foods, we might as truly compare fuels and clothing materials. But there are other reasons for treating them together. These are to be found in the uses to which they are put.

Writers have often compared the action of the body with that of a locomotive, and the analogy is so true in many respects, that we may well follow their example. When the locomotive is doing work by drawing a train, the stoker finds it necessary to feed the boiler fire with constant supplies of coal, and while the coal burns, in other words while its constituents combine with oxygen, a great amount of heat is evolved: this is transmitted to the boiler, and turns large quantities of water into steam. The steam kept under pressure escapes periodically through a valve into the cylinder, forcing the piston out and thus providing the energy required for turning the wheels and pulling the heavy train along. The waste products of the furnace are partly belched forth from the funnel in the shape of steam, carbon dioxide, and unburnt soot, and are partly raked out as clinker, or furnace ash.

Compare this with the progress of food through the body.

During our waking moments we "stoke" our bodies with regular supplies of food. The food passes through slow and complicated processes, but the final result is that, after assimilation in the body, the greater part of the food is really burnt up, *i.e.* oxidised, forming carbonic acid or water, and heat is given out in the process, keeping up the temperature of the body, and the chemical changes involved supply energy for the life

[1] The water vapour can be caught by passing the gases through a tube containing calcium chloride, and the carbonic acid by leading the gases through a vessel containing a strong solution of caustic potash.

[2] Cotton-wool is practically pure cellulose.

processes. In this way is the heat of our bodies supplied, and the energy which is expended in moving our muscles, and in using our brains.

So far the analogy is fairly complete—the stored-up energy of the food, as of the coal, is largely transformed into energy of movement. The locomotive smoke-stack, or funnel, partially represents our breathing apparatus. We exhale from the lungs the products of oxidised (*i.e.* burnt) food-stuffs, and inhale the necessary supply of air. (This latter function is, of course, not common to the locomotive funnel, though air is sucked through the furnace.)

EXPT. 5. **To show the Presence of Carbonic Acid in Exhaled Air.**—Construct an apparatus as shown in Fig. 62 with your smallest flasks, and place some clear lime water in each flask. Place the open end of the T piece in the mouth, and begin to breathe slowly. Owing to the arrangement of the tubes all the inhaled air passes through one bottle, all the exhaled air through the other. Continue the experiment for a few minutes. It will give conclusive evidence that the air breathed from the lungs is rich in carbonic acid. That it is also rich in water vapour you discovered when you first breathed on to a cold window-pane. That it is warmer than the atmosphere under our climatic conditions is also evident. (See also p. 8.)

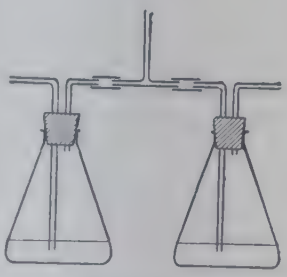

FIG. 62.

We may therefore state that food supplied to the body is oxidised or slowly burnt, and that some of the products which issue from the lungs are the same as those found in the locomotive funnel. In both body and locomotive the material burnt must be constantly replaced.

Respiration is a Combustion.—It was Lavoisier who first clearly stated that the oxygen in the air drawn into the lungs in the act of breathing oxidises, *i.e.* burns, the carbon which has been supplied to the body in food. In 1780 he wrote: "Respiration is a combustion, slow it is true, but otherwise perfectly similar to the combustion of charcoal. It takes place in the interior of the lung. . . . The heat developed by this combustion is communicated to the blood, . . . and is distributed

over the whole animal system." Later he saw that the combustion of hydrogen, which is also supplied in food, takes a part in the production of the heat of the body. As a matter of fact, Lavoisier was mistaken in supposing that it is in the lungs themselves that the union of oxygen with carbon and hydrogen occurs; not till late in the nineteenth century was it recognised that this combination takes place in the tissues of the body.

The Carbonic Acid Gas in the Atmosphere.—Every man or woman, every animal, breathes out carbonic acid gas; and since carbon is present in every fuel, this non-supporter of life and combustion is formed also by every coal fire, by every paraffin lamp, and by every coal-gas flame. And if the air of a room or hall is not changed, the amount of the suffocating gas in the air must continually increase, and at the same time the quantity of oxygen in the air must diminish. Such a condition of the air in a room will always lead to a feeling of oppression and exhaustion or to headache, even if more serious results do not follow, unless the foul air is swept away or changed by a proper system of ventilation.

Then is the amount of carbonic acid gas in the free atmosphere of our earth continually increasing, and is the amount of oxygen continually diminishing, for countless animals are breathing forth carbonic acid gas, and millions of tons of fuel[1] are every year consumed? On the contrary, exact analysis shows that the percentage of oxygen in the open air varies very little, and that the amount of carbonic acid gas remains much the same—rarely more than 0·04 per cent. What, then, becomes of the carbonic acid gas which is pouring into the atmosphere, and how is the stock of oxygen kept up? There must be some counterbalancing action always at work. Can you suggest what this may be?

Let us now consider further, in turn, both fuels and foodstuffs.

FUELS.—During the modern development of industrial arts the use of fuel has enormously increased. Thus the chemical transformation necessary to extract metals from their ores is largely effected by burning coal; the smelting of metals, the manufacture of bricks, pottery, and glass depend

[1] In 1899 Great Britain's output of coal was 220 million tons.

on the intense heat of combustion of fuel: the production of motive power in the steam-engine is brought about by burning fuel. The production of steam for heating and other purposes in the factory also accounts for the combustion of vast quantities of coal. Add to these the domestic consumption; think of the innumerable contrivances for heating houses by coal, coal-gas, wood, or peat, either directly in fire-places and stoves, or indirectly through the medium of hot air, hot water, and steam, and we have enumerated a few of the chief uses of fuel.

A very early, if not the earliest, use of fire was doubtless to cook food. Man has been called the "cooking animal"—and very early in his history, the discovery of fire and its application to cooking vegetable and animal foods gave him an immense advantage over other animals.

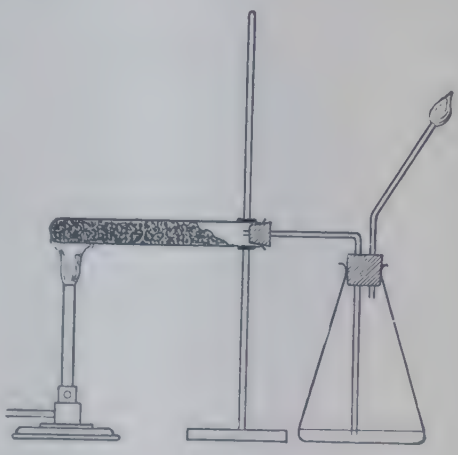

FIG. 63.

Varieties of Fuels.—At first confined to wood, the choice of fuels is now a wide one. The *raw materials*, the various kinds of wood, coal, petroleum, and natural gas, have in part been replaced by its *secondary products*, such as charcoal, coke, coal-gas, methylated spirit, and water-gas.

EXPT. 6. **To find the Effect of heating Coal strongly.**—(1) Weigh out about 10 grams of household coal, broken into small lumps, and place them in a hard glass test-tube. Fix the tube horizontally, and connect it, as shown in Fig. 63, with a conical flask containing a little water. The mouth of the tube leading from the flask is narrowed to form a thick-walled jet. Heat the tube at first gently, and then strongly with a Bunsen burner. When the air is expelled from the flask, and a smoky gas issues, apply a light to the jet. The gas burns with a small luminous flame. It is **coal-gas**. At the same time a black, tarry liquid collects at the end of the tube near the cork. When no more gas is evolved, filter the

contents of the flask ; note the odour of the filtrate, and try its action on litmus paper. It is alkaline.

(2) Crack the tube and examine the residue. It is caked together, hard, black, and brittle. It is **coke**. Place a piece the size of a small bean on a piece of platinum foil, and heat it strongly for a few minutes in a Bunsen flame until it no longer glows. A small grey ash is left.[1] Place it on a watch-glass, and moisten it with water. Does it dissolve? Hardly at all. Add a few drops of dilute hydrochloric acid. Is there any effervescence? Little or none. What substance do you consider the ash contains?

The above experiment represents crudely the **destructive distillation** of coal as it is carried out on an enormous scale in gas retorts, the coal breaking up under the action of heat into *coal-gas*, *coal-tar*, *ammoniacal liquor*, and *coke*. The experiment differs from others we have tried, in the extraordinary complexity of the changes evolved—indeed, so varied are the products, that we cannot profitably study them at this point. To us the experiment is chiefly of interest as showing how coke and coal gas are produced. On the manufacturing scale, however, the ammoniacal liquor which is formed is used for the preparation of ammonia salts, and from the tar is extracted benzene, which is then converted into aniline, the basis of the beautiful coal-tar dyes.

EXPT. 7. **To find the Effect of heating Wood strongly.**—(1) Clean out the flask and tubing used in Experiment 6, take a new hard glass test-tube, and repeat the experiment with weighed pieces of some dry, hard wood. Carefully compare the products obtained from the wood with those which were obtained from coal. Do you not obtain—

(*a*) *Inflammable gases*, which will burn at the jet with a flame not so luminous as that of coal-gas.

(*b*) An acid liquid in the flask ; as a matter of fact, this contains both *wood spirit* and *acetic acid*.

(*c*) A *tarry liquid*, with a curious odour.

(*d*) A residue of *charcoal*, black, brittle, and porous.

(2) Weigh the *charcoal* which is left. It is very light, though its volume is scarcely less than that of the original wood. Also

[1] A known weight of coke may be taken and heated in a crucible in a muffle-furnace, and the weight of ash determined.

heat strongly a weighed portion in an open crucible, and weigh and examine any ash. The very slight grey ash is largely soluble in water ; it effervesces with acids, and is mainly *potash*. A really well burnt piece of wood charcoal consists almost entirely of the element carbon.

Charcoal. — In parts of the Continent, where coal is much more costly than in England, charcoal is largely used as a fuel in stoves. It is obtained in large quantities by setting fire to great heaps of wood built round a central opening almost

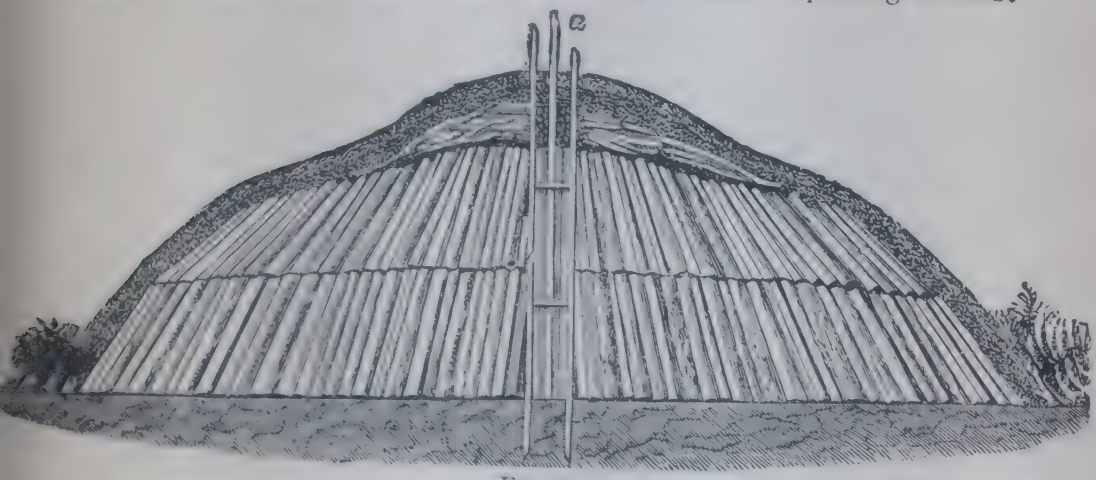

FIG 64

covered with earth so that little air can enter. Some of the wood burns away, but the heat so produced is sufficient to char the rest, the volatile products of course escaping.

EXPT. 8. **To find whether Charcoal is heavier than Water.** — (1) Throw a piece of charcoal into some water. It floats, and seems to be lighter than water. But charcoal has been noticed to be riddled with pores, which of course contain air, so that the piece of charcoal is like a lifeboat with air-compartments. We can get rid of the air, or most of it, in more than one way.

(2) Powder some charcoal finely, and shake it in a test-tube with water. The particles now sink.

(3) Hold a piece of charcoal by a pair of tongs in a flame until it glows, and then plunge it immediately into water. It now sinks.

Charcoal is thus proved to be really heavier than water. Its relative density varies from 1·4 to 1·7.

FOOD-STUFFS.—All the foods which have been examined contain carbon, as well as hydrogen and oxygen. Do they contain no other elements?

It was noticed (see Problem, p. 159) that egg-albumen, when strongly heated, gave off disagreeable fumes in which the presence of ammonia may have been suspected because of its pungent odour. This odour is more evident if albumen is heated with lime, or with soda-lime, which is quicklime slaked with caustic soda.

EXPT. 9. **To observe the Effect of Heat on Egg-Albumen mixed with Soda-lime.**—Mix a very small quantity of dried, powdered egg-albumen with five or six times as much dry, powdered soda lime, place the mixture in a small test-tube, and add a little more soda-lime. Heat very gently, and smell cautiously the mouth of the tube. The odour of ammonia is easily detected. Also hold a moist red litmus paper in the mouth of the tube: the paper is at once turned blue.

Hence, since albumen to a less degree also gives off the odour of ammonia when heated alone, it is evident that **albumen contains the "spirit of sal-ammoniac"** or **its constituent elements**, whatever they may be.

PROBLEMS

Find the effect of heating the following with soda-lime in dry tubes. Classify them according as ammonia is or is not evolved.

Pure starch	Wheat flour	Olive oil
Lentils (powdered)	Gluten[1]	Mutton fat
Peas (,,)	Dry lean meat	Cheese

THE COMPOSITION OF AMMONIA.—The discovery that certain food-stuffs, *e.g.* egg-albumen, evolve ammonia when heated with soda-lime, raises the question of the composition of this substance. Is it a distinct kind of gas? Is it an element like oxygen, or a compound like carbonic acid gas? The answer may show that certain food-stuffs contain some other important element or elements besides carbon, hydrogen, and oxygen.

It has already been found (Vol. I., p. 160) that when sal-ammoniac is heated with quicklime the "spirit of sal-am-

[1] Gluten can be obtained from flour by enclosing some flour in a muslin bag in water, and squeezing. Starch passes through, and leaves gluten.

moniac," usually simply called "ammonia," is evolved, which is extremely soluble in water. Can this substance be collected as a gas? It is clearly useless to attempt to collect it over water. Let us try to collect it by upward displacement.

EXPT. 10. **To attempt the Collection of Ammonia.**—Mix a little sal-ammoniac with twice its volume of quicklime, and place the mixture in a dry test-tube, fitted with a cork and tube bent at right angles. Fix the tube in a clamp as shown in Fig. 65. Place a dry test-tube over the delivery tube, and waft a Bunsen burner underneath the mixture. Presently a very pungent odour can be detected at the mouth of the vertical tube, and the fumes will turn a moistened red litmus paper blue. Lift the test-tube gently up, close its mouth with the thumb, and place another test-tube over the delivery tube.

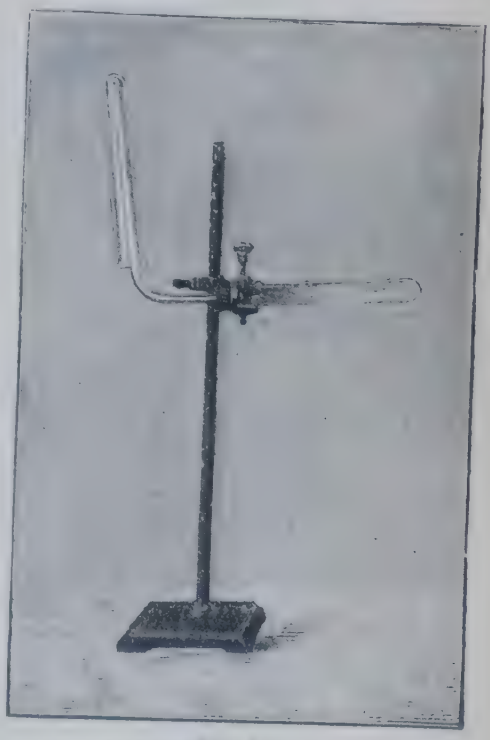

FIG. 65.

To find out whether a gas has been collected, and if so what some of its characters may be, try the following simple experiments:—

(1) Place the mouth of the tube, closed by the thumb, below the surface of some water, remove the thumb, and shake gently. The water rises gradually up the tube. (If all the air has been expelled from the apparatus, the water will completely fill the tube.) This result shows that a gas has been collected, and that **it is very soluble in water.** As a matter of fact, 1 c.c. of water at the ordinary temperature will dissolve about 800 c.c. of ammonia gas.

(2) Hold a tube full of the gas mouth upwards for one or two minutes, and then find whether it remains full. On the contrary, all, or nearly all, has escaped. Try also to pour the gas *upwards* from one tube to another. The results obtained show that

ammonia gas **is much lighter than air**; its relative density is 0·59 (air = 1).

(3). Fill a dry jar with the gas, and plunge into it a lighted splint of wood. The flame is extinguished, but observe that the *gas itself seems to have a tendency to burn.*

(4) The gas has also been observed to possess a very **pungent odour**, and to have an **alkaline character**.

These experiments show clearly that the "spirit of sal-ammoniac" can be collected, and that it is a distinct kind of gas. It was first collected by Dr. Priestley in 1774; he collected it

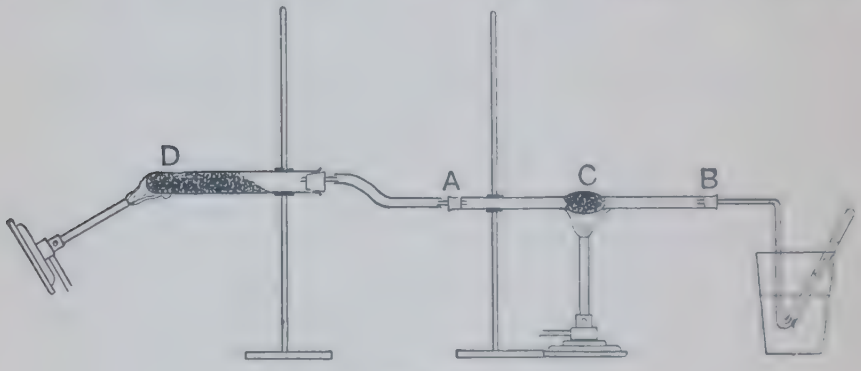

FIG. 66.

over dry mercury, and called it "*alkaline air*," but we shal retain the name ammonia.

Let us now proceed to consider whether ammonia is an elementary gas, or whether it is compound like carbonic acid gas. The only fact noticed thus far, which makes it somewhat doubtful whether ammonia is a simple gas, is the observation that it seems inclined to burn with a brown flame. Does it contain some inflammable gas, either hydrogen or some other? Will ammonia under more favourable conditions unite wholly or in part with oxygen; and if so, what are the products?

EXPT. 11. **To find whether Ammonia can be oxidised by Copper Oxide.**—Pack the bulb C of a "reduction tube" with dry granular copper oxide, and connect it, as shown in Fig. 66, with a test-tube, D, from which the ammonia is expelled.[1] Have ready also a short delivery tube, so that any gas which

[1] Ammonia gas can also be conveniently prepared by warming 4 parts of sal-ammoniac with 3 parts of slaked lime, and a little water, in a flask. The gas should then be dried by sending it through a tube packed with small lumps of quicklime.

passes through the tube may be collected in test-tubes over water. Warm the tube D gently (do not yet connect the delivery tube with the end B), and as soon as there is evidence of ammonia escaping at B, begin to heat the bulb C with another flame, at first gently, but then strongly. Notice that a dew begins to be deposited in the tube near B, and that the escape of ammonia finally ceases. Then connect the delivery tube, and collect the gas which escapes over water in test-tubes. Notice also that the copper oxide becomes red, and is clearly being reduced.

What is the gas collected? Test it with a burning splint and with lime water. It does not support combustion, it has no odour, it does not turn lime water milky. It has, in fact, all the characters of **nitrogen**. The dew deposited in the tube appears to be **water**: prove that this is so.

This experiment proves that ammonia must contain nitrogen and hydrogen, and perhaps nothing else. This is indeed the case: more complete experiments, which we cannot stay to perform, prove that **Ammonia is a compound of nitrogen and hydrogen**. Priestley[1] made an experiment similar to the above, but he made use of oxides of lead or of iron: he noticed the liberation of the gas which we call nitrogen, and the formation of water, and also how the oxide was "revived," as he expressed it.

Nitrogenous Food-Stuffs.—Since certain food-stuffs, *e.g.* egg-albumen, evolve ammonia when heated alone or with soda-lime (see Experiment 9), the discovery that ammonia contains nitrogen leads to the important conclusion that *these food-stuffs contain nitrogen* in addition to carbon, hydrogen, and oxygen. Nitrogen is, in fact, contained in the white of egg, in gluten, which is a constituent of wheat-flour, in milk, and in the muscular fibre of lean meat. To distinguish these and others from food-stuffs, such as sugar, which do not contain nitrogen, they are often called **nitrogenous food-stuffs**.[2]

Organic Foods may now be classified as follows:—

1. *The Carbohydrates* — including all sugars, starch, and gummy materials.

[1] Priestley's *Natural Philosophy*, vol. iii. p. 189.

[2] They are also called *proteids*. It may be added that some of them also contain a small amount of sulphur and phosphorus.

2. *The Fats*—including not only solid fats, but also semi-liquid butters, and many liquid oils, all of which can be proved to be similar in their chemical composition.

3. *The Nitrogenous Food-stuffs*—including lean meat, egg-albumen, the gluten of wheat-flour, and the casein of milk.

The human body in its muscular fibre, in its nerve tissues, in the blood and elsewhere, contains nitrogenous materials, and it is therefore evident that if a healthy, vigorous body is to be maintained, the diet must include nitrogenous food-stuffs, in addition to carbohydrates and fats, which contain no nitrogen. It might be supposed that the addition to the diet of a little nitre or sal-ammoniac or some other inorganic store of nitrogen might provide the necessary nitrogen; but experience shows that we cannot in our systems build up from such a source the nitrogenous fabric of our bodies. Our bodies must be supplied with the nitrogen in the form of some organic material. Vegetarians may find this necessary nitrogen in lentils, beans, and peas, but most people in this country obtain it mainly in the animal form of lean meat.

In addition to the classes of foods already spoken of, there are others which are essential if health is to be maintained. Certain other elements must be provided, which may be found in common salt, in potash, and in other **mineral "salts."** These salts are present in larger quantity in fruits and vegetables than in other foods, and hence it is that a proper proportion of fruits and vegetables should be present in any scheme of diet. Large quantities of **water** must of course also be drunk.

Animal tissues contain nitrogen, and when they putrefy and decay ammonia is liberated: it is therefore found in sewage, and hence, if a water is found to contain more than a very small amount of ammonia, it may be suspected that it is contaminated, and it should be considered quite unfit for drinking purposes.

Problems

1. Find whether dry ammonia gas can be decomposed by heat alone. Try by passing the gas, dried by lime, through a bulb tube packed

with clean copper turnings (on which ammonia has no action). Collect the gas which issues over water, and find whether it is inflammable.

2. Find whether ammonia gas will burn when mixed with oxygen. Try by bubbling oxygen slowly through a strong solution of ammonia warmed gently in a beaker, and apply a flame to the mouth of the beaker.

3. Find what occurs when a hot platinum wire bent into a spiral is hung in the mixture of oxygen and ammonia gas as described in No. 2.

4. Find what occurs if common air is bubbled through a strong solution of ammonia, and then over metallic copper heated in a bulb tube.

5. Examine the effect of heat upon the substance provided. What elements do you find it to contain?

CHAPTER XVI

THE TRANSFORMATION OF MATTER.—THE LAW OF DEFINITE PROPORTIONS

ONE great fact which has stood out in every inquiry and research upon which we have entered is this, that matter can undergo change or transformation. If we consider our work more closely, we shall find that many of these very numerous instances of the change of matter can be classed under three heads :—

Decomposition.—Not infrequently the effect of heat has been to produce two substances from one. Thus :—

Chalk = quicklime + carbonic acid gas.
Red lead = litharge + oxygen.
Water = hydrogen + oxygen.

Such a change is a *decomposition* or an *analysis*.

Combination. Numerous instances of combination were met with in our study of Fire and Air. Thus :—

Lead + oxygen = litharge, or else red lead.
Magnesium + oxygen = magnesium oxide.
Phosphorus + oxygen = oxide of phosphorus.
Hydrogen + oxygen = water.

In contrast to analysis a change of this sort is called a *synthesis*.

Decomposition and Simultaneous Combination. We have had frequent instances of more complicated changes than simple decompositions or combinations. Thus, when water vapour was passed over red-hot iron, the water was

decomposed, hydrogen being set free, but at the same time iron was oxidised and iron oxide formed. Or,

Iron + water = hydrogen + iron oxide.

Other instances of decomposition associated with combination were :—

Hydrogen + copper oxide = copper + water.
Iron + nitric oxide = nitrogen + iron oxide.

Many other changes were only partially studied, and cannot as yet be classified with certainty. Try, for instance, how far you can complete the following equations as statements of the whole of the change effected :—

Lime water + sodium carbonate =
Calcium chloride + sodium carbonate =
Zinc + sulphuric acid =
Copper + nitric acid =
Nitre + sulphuric acid =
Hydrochloric acid + caustic soda =

Although much time has been spent, in the preceding pages, in examining these and other changes, yet in very few cases has the whole course of the change been verified, and it is therefore premature to classify them definitely.

All the varied forms of change instanced above are what are called **chemical changes**, in distinction from those physical changes, of solid into liquid, or liquid into vapour, and the reverse, which have already been considered. Some of the characteristics of chemical change will be brought out in the succeeding pages.

The changes which matter can undergo are thus of the most varied character, and though matter may become invisible, formless, and vague, yet we assured ourselves that *matter was not destroyed* in several test cases (*see* p. 50). We therefore assumed again and again in our experiments that the sum of the masses of the substances formed as the result of a change was equal to the sum of the masses of the substances taken. As Lavoisier said, "C'est sur ce principe qu'est fondé tout l'art des expériences en Chimie"—the whole art of making experiments in Chemistry is founded upon this principle.

All substances may be classified into two divisions, as homogeneous or heterogeneous, according as any portion appears or does not appear, so far as our powers of observation go, to be of the same kind throughout, however minutely it is examined.

(*a*) **Homogeneous.**[1]—Such are oxygen, chalk, water, solder, a solution of salt in water, air, &c. These will be considered in detail directly.

(*b*) **Heterogeneous.**[2]—For example, a plum pudding, blood, or a mixture of sand and nitre. Plum pudding is compounded of suet, raisins, currants, lemon-peel, flour, &c., in proportions which may be varied at the discretion of the cook, but some at least of the constituents—the currants, the raisins, and the lemon-peel—may be distinguished and picked out from the finished pudding. If a drop of blood is examined microscopically, it is seen to contain minute solid discs or corpuscles of two kinds floating in a clear liquid, the one pale red in colour, the other colourless or white. So too in a mixture of sand and nitre, however intimately they are mixed together, yellow grains of sand and tiny white crystals of nitre can be readily distinguished when the mixture is examined under a microscope. If a little of the mixture is placed in the mouth, the nitre is dissolved with its characteristic taste, and the hard gritty insoluble sand is left upon the tongue. It is evident that the sand and nitre are still present as such.[3]

Chemistry deals as a rule only with homogeneous substances: if, however, a chemist has to deal with a heterogeneous substance, he considers its homogeneous constituents separately.

Homogeneous substances can themselves be classified under three heads as—

(1) Elements.
(2) Chemical compounds.
(3) Indefinite compounds.

1. ELEMENTS.—Many homogeneous substances can undergo change by decomposition and be separated into different

[1] Greek: *homos*, the same, and *genos*, kind.
[2] Greek: *heteros*, other, different, and *genos*, kind.
[3] Heterogeneous substances, such as a mixture of sand and nitre, are often called "mechanical mixtures." The expression is, however, open to serious objection, as the only meaning to be attached to the term "mechanical" is "chemical," and requires the definition of the term "chemical."

substances. Chalk will yield an equal weight of quicklime and carbonic acid gas. Carbonic acid itself can be decomposed, and we know it to contain both oxygen and carbon. Water can be decomposed into oxygen and hydrogen. How about carbon, or oxygen, or hydrogen? These substances can combine with others, but we have had no evidence of their decomposition into anything else. So far as we know, the decomposition can go no further. Such a substance we have termed an *element*, and we have defined the term thus :—

An element is a substance which has not by any known process been separated into two or more different substances.

2. CHEMICAL COMPOUNDS.—The term compound has been freely used in previous chapters of this book, though no statement has been made of the precise meaning to be attached to it. For instance, it has been stated that

> Chalk is a compound of quicklime and carbonic acid,
> Magnesium oxide is a compound of magnesium and oxygen,
> Water is a compound of hydrogen and oxygen.

Let us consider these substances in detail. They all differ, in the first place, from a mixture of sand and nitre in being *homogeneous*.

Secondly, in them the **characteristic properties of those substances of which they are composed have disappeared, and other properties are found.** The unaided senses will never suggest that water is capable of separation into two invisible gases, and will find no evidence of the metal magnesium in the white infusible earth magnesia, nor of carbonic acid gas in chalk or marble. On the contrary, chalk, magnesium oxide, and water all possess distinctive properties which are markedly different from those of their respective constituents.

In the third place, a remarkable **definiteness of composition** has been found, *e.g.* :—

(1) *Chalk.*—The most careful analyses that we have made with powdered calc-spar show that 100 grams of it always yield very nearly 56 grams of quicklime and 44 grams of carbonic

acid (*see* p. 143). We have also combined quicklime with carbonic acid by different processes, and found that 56 grams of quicklime will in each case yield a white earth-like body resembling powdered calc-spar, which when dried weighs 100 grams; each product is a compound of 56 grams of quicklime with 44 grams of carbonic acid. These products, therefore, possess the same quantitative composition, while being also alike in other respects, and the same name "chalk" may be given to each. Chalk, then, is a substance which possesses this definite composition.

The more care experimenters have taken in studying this question the more nearly have their own experiments agreed with each other and with those of other experimenters. We, therefore, may conclude that the small differences which do exist between results point not so much to innate differences in composition as to *errors of experiment* which even the most experienced chemist cannot wholly avoid.[1] So satisfied are we of this definiteness of composition, that if a sample of "chalk" were given us, which on careful analysis gave different results, we should certainly conclude that it was not chalk. What then? Either some other substance or else impure chalk—that is to say, chalk mixed with something else.

(2) *Water.*—We studied the composition of water in detail, and arrived at these results:—

Water, when decomposed by the electric current, produced 2 volumes of hydrogen for every 1 volume of oxygen.

2 volumes of hydrogen and 1 volume of oxygen combined and produced water.

9 parts by weight of water are composed of 8 parts by weight of oxygen, and 1 part by weight of hydrogen.

Careful experiments with the substance we call "water" will always lead to the same results, whether it be a sample of rain water which has just descended from the clouds, or it has been formed by burning a jet of hydrogen, or it has been obtained by distillation from the Atlantic Ocean.

[1] This kind of evidence must, of course, be accepted cautiously, and many of the most remarkable discoveries of science have resulted from the close scrutiny of small "errors." We need only instance the discovery of Neptune, from small differences between the observed and calculated positions of Uranus; and the difference in density between different samples of "nitrogen" which has led to the discovery of the gas "argon" in the atmosphere.

In contrast to the definiteness of composition of these substances is the composition of heterogeneous substances, such as a plum pudding, blood, or a mixture of sand and nitre. Supposing different portions of *one* particular mixture of sand and nitre were carefully analysed several times, the same results would of course be obtained. Other mixtures of sand and nitre could, however, be made, each of which would be called "sand and nitre," but their composition would not be the same. Again, samples of blood differ in composition. So also puddings, each called "a plum pudding," may and almost certainly will differ in composition.

We have, then, in the course of our study met with various substances, of which chalk and water are instances, to which the following statements apply :—

(1) They are homogeneous substances containing two or more elements.

(2) Some of their properties are not the same as those of their constituents, nor even a mean of those of their constituents.

(3) Two samples even when produced by different processes and by different experimenters, or obtained from different sources, possess the same quantitative composition while being alike in other respects.

The statement that there are substances of which this is true may be called the "**Law of Definite Proportions.**"

Substances to which these three statements apply are called "**definite**" or "**chemical compounds.**"

Any change resulting in the production or decomposition of a chemical compound is a "chemical change," and a chemical change is therefore associated with those features which are proper to a chemical compound.

PROBLEMS

Obtain samples of the following substances from different sources or makers, and find whether they are definite or indefinite compounds : —

1. *Potassium Chlorate.*—Weigh a dry hard glass test tube. Introduce about 2 grams of dry potassium chlorate, and with a camel's hair brush dust all particles down to the bottom of the tube, and again weigh.

Fix the tube in a clamp, and heat gently both the bottom of the tube and the part just above the substance; finally heat with a full flame. Allow to cool and then weigh. Calculate the percentage loss of oxygen.

One sample was found to contain 39 per cent. of oxygen. What do you find?

2. *Red Oxide of Mercury.*—Similarly heat in a weighed hard glass test-tube quite 2 grams of dry red oxide of mercury, but only heat the bottom of the tube.

A sample was found to contain 7·5 per cent. of oxygen. What do you find?

3. *Oxide of Copper.*—Weigh about 1 gram in a porcelain boat, and place in a hard glass tube, and arrange to pass a current of dry hydrogen[1] through the tube. After expelling all the air, begin to heat the tube, and continue until the oxide is apparently completely reduced to the red metal. Allow to cool in a current of the gas, and then carefully pull out the boat with a hooked copper wire and weigh it. Reheat and reweigh. Calculate the percentage loss of oxygen.

A sample was found to contain 20·3 per cent. What do you find?

4. *Sodium Carbonate* (anhydrous).—Determine the percentage weight of carbonic acid gas which can be expelled by acid. Proceed as in Experiment 1, p. 140).

A sample contained 41·5 per cent. What do you find?

3. INDEFINITE COMPOUNDS.—There are other homogeneous substances containing more than one element which are not characterised by that definiteness of composition to which we have referred. Such are solutions, alloys, and mixed gases.

Solutions of salt in water, for example, each called "a salt solution," may vary greatly in composition. Moreover, to a given solution we can within certain limits add either salt or water, and we shall obtain a fresh solution of salt in water, which will still be homogeneous.

Solder, again, is an alloy of tin and lead. We have not ourselves analysed different samples of solder, but if we did we should find that, though homogeneous and closely alike in other respects, yet too great differences would exist between the ratios of lead to tin in different samples of solder to be due to any error

[1] Coal-gas may be used in place of hydrogen

of experiment. We may say that solder is an *indefinite* compound of tin and lead.

Is *common air* a chemical compound of oxygen and nitrogen? In the first place we recall that substances will burn in oxygen as well as in common air, but much more brilliantly, and that oxygen supports the life of animals, as does common air, but more vigorously; in fact, the properties of common air are a mean of those of oxygen and of the inert gas nitrogen. But do samples of air collected at different times and places not possess the same quantitative composition? Very exact experiments have recently been made by Leduc. He determined the weight of oxygen in the air, and found—

At Paris	Summer ...	23·23	per cent. of oxygen.
,, Dieppe	July	23·16	,, ,,
,, ,, ...	April ...	23·07	,, ,,
On Belgian frontier ...	Summer ...	23·17	,, ,,
,, ,, ,, ...	Winter ...	23·09	,, ,,

So accurate were his experiments, that the differences between these numbers cannot be put down to experimental errors. They show that there are distinct though slight variations in the composition of the air. On these grounds common air cannot be regarded as a chemical compound.[1]

The Existence of an Element in the Separate and in the Combined State.—We have stated that magnesium oxide is a chemical compound of magnesium and oxygen. We do not mean, however, that the white infusible earth actually contains magnesium as a *metal*, and oxygen as a *gas*. That is inconceivable. What we do mean is that magnesium oxide contains the material or substance of these elements. A similar statement is true of other chemical compounds. Water, for instance, does not contain oxygen or hydrogen as gases, but the substance of these elements. Why, when elements are thus intimately compounded, their characteristic properties should vanish, and why other properties should appear, is, at present, an unanswered question.

1 We do not suppose that any element of chemical and indefinite compounds is complex of eleven ruthenium. On the contrary, we believe that at present it is impossible to define the one clearly from the other. This is shown by the fact that there are some substances with regard to the classification of which as chemical or indefinite compounds there is wide difference of opinion amongst even trained chemists.

Hence, to be accurate, a statement of the composition of a chemical compound is merely the statement of its formation or of the changes of which it is capable.

Pure Materials.—It is not easy to prepare substances *pure*, that is, free from admixture with other substances. Adulteration Acts have made it easier for manufacturers to be honest, but it is still difficult to prepare pure materials, as may be seen by quoting from a recent price list :—

		lb.	oz.
Acid hydrochloric ...	commercial	3*d.*	
	pure	6*d.*	
	pure for analysis	1*s.*	
Soda caustic	lump	8*d.*	
	white stick	2*s.*	3*d.*
	pure	...	6*d.*
	from sodium	...	1*s.* 6*d.*
Zinc	granulated	6*d.*	
	free from arsenic	2*s.*	
	distilled, free from iron ...	6*s.*	

Each of the above substances is homogeneous : the first is an indefinite compound, the second a definite or chemical compound, and the third an element. It is clear, however, that the commercial article may in each case contain foreign matter.

Boyle in his day did not always find that the materials with which he was supplied were pure. In *The First Essay on the Unsuccessfulness of Experiments*, he speaks of seed of clover grass from a great quantity of which no grass sprang up, and of varying samples of vinegar which did not all contain living creatures, and then of his own difficulties, thus :—

"The other day, having occasion to use some *Spirit of Salt*, whereof I was not then provided, I sent for some to a Chymist, who making it himself, was the likelier to afford that which was well made ; but though I gave him his own rate for it, at the first Rectification even in a Retort, a single pound afforded us no less than six ounces of phlegm ; and afterwards being further rectified in a high body and gentle heat, the remaining Spirit parted with a scarce credible quantity of the like nauseous liquor, and after all these sequestrations of phlegm was not pure

enough to perform what we expected from it. Of which complaining to an excellent Chymist of my acquaintance, he sent for Spirit of Salt to a very eminent distiller of it, who gets much by his profession. and passeth for a very honest man: but this spirit, besides its weakness, discover'd itself to be sophisticated with either Spirit of *Nitre*, or *Aqua fortis*, which betray'd itself by its peculiar and odious smell; whereas Spirit of Salt skilfully and sincerely drawn, is commonly of a greenish colour, bordering upon yellow, and hath usually a Peculiar, and sometimes (as I can exemplifie to you in some of mine) a not Unpleasing smell."

The greenish colour which Boyle regarded as a test of purity we should now think a proof of impurity due to the presence of iron.

We too must not rashly assume that the labels on all bottles truly denote their contents. For instance, sulphuric acid contains a small percentage of water, zinc usually contains traces of other metals and of carbon, caustic soda is rarely free from carbonate of soda, and common salt often has other salts mixed with it.

Exercises

1. Try whether samples of zinc and iron will dissolve entirely in dilute sulphuric acid.

2. Try whether samples of alcohol, sal-ammoniac, and distilled water will evaporate completely without leaving any residue.

3. Endeavour to prepare pure samples of the following commercial substances:—

a. Copper Sulphate, by recrystallisation.

b. Common Salt, by passing into a saturated solution of the salt the fumes of the "spirit of salt," just as the fumes were in Vol. I, p. 156, Expt. 5, passed into water. Crystals of pure salt will be thrown out of solution. When a considerable quantity has separated out, collect the salt on a filter, wash it with a little pure strong hydrochloric acid, drain it by a Bunsen pump, and lastly heat it gently in a porcelain basin.

PUBLIC EXAMINATION PAPERS

OXFORD LOCAL EXAMINATIONS

Junior Candidates

ELEMENTARY SCIENCE. Paper II

July 1900

2 Hours

1. What are the gases contained in air, and in what proportions are they present? Describe some experiments which demonstrate their existence, and also their most characteristic properties.

2. Briefly describe three methods of preparing hydrogen from water. What volume of hydrogen could be obtained from 10 grams of water (1 litre of hydrogen weighs 0·09 gm.)?

3. What happens if one passes a stream of hydrochloric acid gas through lime water?

If the amount of gas absorbed by the solution is known, it is possible to calculate in what quantities certain other substances are formed. Why is this, and what important principle is thereby illustrated?

4. To what is the hardness of water due? When and how is it possible to remove this hardness?

Describe very briefly the chemical properties of the substances which make water hard.

5. What takes place when a solution of potash is neutralised by adding nitric acid? Describe one or two simple experiments to demonstrate the nature of the body formed.

6. What are the chief bodies present in coal-gas? What are the substances formed when it is burnt in the ordinary gas flame? How would you demonstrate their presence?

7. What are the metals gold, iron, zinc, and lead most commonly used for, and why?

CAMBRIDGE LOCAL EXAMINATIONS

Preliminary Examinations

CHEMISTRY

December 1900

1 Hour

1. Give a short account of the chief constituents of the air. How is the relative amount of each constituent affected (*a*) by plants, (*b*) by animals?

2. Describe the preparation and chief properties of hydrogen.

3. How would you prepare hydrochloric acid? Give *two* tests by which a solution of hydrochloric acid can be distinguished from a solution of nitric acid.

4. What are the principal properties of carbonic acid gas? What is its action with (*a*) lime water, (*b*) a solution of litmus?

5. Explain carefully what is meant by *an element*. How would you show that coal-gas contains carbon and hydrogen?

CAMBRIDGE LOCAL EXAMINATIONS

Junior Candidates

ELEMENTARY EXPERIMENTAL SCIENCE. Paper II [1]

December 1900

4. Describe and compare the physical and chemical properties of any three common solid elements and of one compound of each.

5. Describe a method of preparing oxygen. How would you prove that oxygen is a constituent of (i) air, (ii) chalk?

6. Give an account of the physical properties of water.

7. Describe carefully how you would determine the composition of water by weight, giving drawings of the apparatus.

[1] The syllabus has very recently (1902) been revised, and, as it now stands, is completely covered by the course described in this book. The examination in 1900 was upon an old syllabus, but the questions quoted above are upon subjects scheduled in the new syllabus.

JOINT SCHOLARSHIPS BOARD

Intermediate Scholarships Examination

ELEMENTARY DRAWING AND ELEMENTARY EXPERIMENTAL SCIENCE[1]

June 1900

3 Hours

3. Measure the area of this examination paper in square centimetres, and from the number obtained calculate the length in inches of a square sheet of paper having exactly the same area (1 metre = 39·37 inches).

4. Describe the appearance and obvious properties of the following substances: (1) a lump of chalk; (2) a block of ice on a summer's day; (3) a stick of brimstone.

5. The following observations of a barometer were made at noon each day during the first two weeks in February:—

Feb. 1 ...	750 mm.	Feb. 6 ...	762 mm.	Feb. 11 ...	758 mm.
,, 2 ...	754 ,,	,, 7 ...	738 ,,	,, 12 ...	756 ,,
,, 3 ...	756 ,,	,, 8 ...	738 ,,	,, 13 ...	755 ,,
,, 4 ...	757 ,,	,, 9 ...	745 ,,	,, 14 ...	750 ,,
,, 5 ...	760 ,,	,, 10 ...	750 ,,		

Draw a curve representing the variations in atmospheric pressure, and briefly comment on the kind of weather likely to have been prevalent during the period under review.

7. Sketch and describe an apparatus such as you would use if you were requested to ascertain if the constituent of air remaining after the rusting of iron in a confined volume of air produces any change when passed over heated copper.

8. Give an account of experiments you could make if you were required to show with the aid of a balance that the gas evolved on pouring an acid on chalk is identical with the gas lost during the conversion of chalk into lime. Describe a lime-kiln.

[1] The questions on Elementary Experimental Science alone are quoted here. The syllabus of this part of the examination is in accordance with the Elementary Science Syllabus of the Incorporated Association of Headmasters. "Physics," 88 pp., "Chemistry," 88 pp., 1s. (Educational Supply Association, 41, Holborn Viaduct, E.C., price 3*d*.)

JOINT SCHOLARSHIPS BOARD

Intermediate Scholarships Examination

ADVANCED CHEMISTRY[1]

June 1900

1½ Hours

You are not expected to answer more than 4 *questions.*

1. Describe experiments from which it may be inferred that air consists chiefly of two gases present in the ratio of about 4 : 1.

2. When lead is heated in air the substance litharge is formed. How would you proceed to show that the production of this substance is accompanied by a change in weight, and that this change is due to the absorption by the metal of some of the active constituents of the air?

3. Describe an experiment which points to the conclusion that when coal-gas burns in air we may equally well regard the combustion as that of air in coal-gas.

4. When acid is poured on chalk an invisible gas which does not support combustion is evolved. Give an account of experiments you would make if you were required to determine if this gas is or is not identical with the inactive constituent, comprising four-fifths by volume, of the air.

5. A glass of muddy water is given you. How would you ascertain if it contained dissolved solid matter, and how would you find the weight of dissolved solid matter in 100 c.c. of the water?

6. Give experimental evidence that such substances as flour, sugar, and alcohol contain carbon.

JOINT SCHOLARSHIPS BOARD

Major Scholarships Examination

ELEMENTARY SCIENCE (Physics and Chemistry)[2]

July 1900

2 Hours

1. (a) Give a full description and sketch of the apparatus you would use if you were required to determine the volume in cubic centimetres of a large, irregular-shaped flint.

[1] The syllabus is that of the L.A.H.M. Elementary Science Syllabus, "Chemistry."

[2] The syllabus of the *Chemistry* is that of the L.A.H.M. Elementary Science Syllabus, "Chemistry."

(*b*) The density of silver referred to water is 10·26; the density of lead referred to silver is 1·075. Assuming that when lead and silver are alloyed no change in volume occurs, find the percentage composition of a silver-lead button from the following data :—

Weight of button in air, 15·167 grams.
Weight of button in water, 13·773 grams.

2. On the squared paper provided, and with the aid of the table given below, draw a curve representing the duration of night from March 1st to June 29th. Assume that day commences half-an-hour before sunrise and ends half-an-hour after sunset.

	Sun rises.	Sun sets.
Mar. 1st	6.48 A.M.	5·37 P.M.
,, 11th	6.26 ,,	5·55 ,,
,, 21st	6. 3 ,,	6.12 ,,
,, 31st	5.40 ,,	6.29 ,,
Apr. 10th	5.18 ,,	6.45 ,,
,, 20th	4.56 ,,	7. 2 ,,
,, 30th	4.36 ,,	7.18 ,,
May 10th	4.18 ,,	7·34 ,,
,, 20th	4. 3 ,,	7.49 ,,
,, 30th	3·52 ,,	8. 3 ,,
June 9th	3·45 ,,	8.13 ,,
,, 19th	3·44 ,,	8.18 ,,
,, 29th	3·48 ,,	8.18 ,,

3. Describe a barometer and explain its use. What experiments would you make if you were required to demonstrate that the change in level of the mercury in a barometer depends on variation in atmospheric pressure?

4. Describe some well-known facts which show that substances usually expand when heated. How would you experimentally determine which of the metals iron and copper is the better conductor of heat?

5. Mention some of the physical differences between sea water and fresh water. How would you proceed to discover if sea water contains dissolved air, and how would you collect a sample of such air and determine if its composition is identical with that of ordinary air?

6. Describe the appearance and obvious properties of copper, sulphur, gypsum, flint, and litharge.

7. A black powder may or may not be copper scale. How would you experimentally decide the point?

8. Describe experiments which show that chalk when burnt is greatly changed.

Make a rough sketch of a lime kiln to illustrate your answer.

UNIVERSITY OF LONDON

Matriculation Examination

GENERAL ELEMENTARY SCIENCE

The following are the Questions on Chemistry in the papers set January 1900 to June 1901

January 1900

1. You are given a number of cylinders containing ammonia gas. Describe the experiments you would make in order to illustrate the chemical and physical properties of the gas.

2. Describe any experiments showing that the chemical combination of two elements to form a definite compound takes place in a fixed proportion by weight.

3. What is the effect of heating (*a*) chalk; (*b*) rock-crystal; (*c*) a mixture of hydrogen and oxygen; (*d*) carbon in air; (*e*) sulphur in air? Name and give the formula of each substance produced.

4. Enumerate the differences in physical properties exhibited by sea and fresh water respectively. How can a specimen of common salt be prepared from sea water?

5. How would you separate from each other a mixture of sand and chalk?

6. Describe an experiment showing that water and carbon dioxide are formed in the combustion of a candle. How can ordinary coal-gas be made to burn with a non-luminous flame?

June 1900

7. The words "combustion," "flame," and "incandescence" occur in the syllabus. Give instances of each, and state how these phenomena differ from each other.

8. What changes occur when ice is heated to 0° C.? When it is heated above 0°? When water is raised to 4°? When it is raised to 100°? When steam is heated to 200°?

9. When an ordinary wax-match is "struck" it takes fire, and the wax is ignited. Explain the physics and chemistry of the changes which occur.

10. Gunpowder is made by mixing together charcoal, sulphur, and saltpetre: describe fully how you would propose to obtain each of these ingredients in a state of purity from a specimen of powder.

11. A natural water is found to give a white precipitate (*a*) on boiling, (*b*) on the addition of a solution of silver nitrate, (*c*) on the addition

of a solution of barium chloride: what inferences would you draw from this behaviour, and how would you test their accuracy?

12. Describe those properties possessed by sulphuric, hydrochloric, and nitric acids, which lead to these substances being designated "acids."

January 1901

13. How could a diamond, a piece of rock-crystal, a piece of calc-spar be distinguished from each other?

14. State what you know about carbon dioxide.

15. Describe how ammonium nitrate can be prepared. How can ammonia and nitric acid be obtained from it?

16. Describe the experiments which could be performed in order to demonstrate the points of difference between an aqueous solution of ammonia and of caustic potash respectively.

17. Describe and explain the various methods with which you are acquainted by means of which air can be deprived of its oxygen.

18. Common salt is dissolved in distilled water. How are the freezing point, boiling point, and temperature of maximum density affected thereby? Draw and describe the apparatus with which you would propose to determine the freezing and boiling points of water?

June 1901

19. Why does smoke go up a chimney? Give a detailed explanation of this.

20. What is formed when phosphorus burns in air? Can it be made to burn in any other gas? and if so, in what?

21. What is left behind as a gas when phosphorus burns in a closed vessel? What properties has it? Has it been made to combine with oxygen, and if so, how?

22. Common salt, sulphur trioxide, carbon dioxide, and metallic sodium dissolve in water; ammonia, chalk, and metallic zinc dissolve in dilute hydrochloric acid; sulphur dissolves in carbon bisulphide: state what in each case remains on evaporating off the solvent.

23. How can chlorine be obtained from common salt? Sketch the apparatus you would employ for its preparation, and describe three experiments illustrating its properties.

24. How can carbon, sulphur, phosphorus, and chlorine respectively be made to combine with hydrogen? Give the names of the compounds formed in each case.

SOME PUBLIC PRACTICAL EXAMINATIONS

OXFORD LOCAL EXAMINATIONS

Junior Candidates

ELEMENTARY SCIENCE. Paper III[1]

July 1900

2 Hours

Write a full account of all your work, and make sketches of the apparatus you use.

[Attempt 2 questions, one from each section. You must try either question 1 or question 3, but are not forbidden to try both.]

1. Determine the specific gravity of the two solids, P and Q,[2] provided.

2. Using only a test-tube of water, a thick copper wire, and a thermometer, perform and describe experiments to illustrate the various ways in which heat may be transferred from one place to another.

3. Collect over water some of the gas evolved on heating the substance R with strong sulphuric acid. Determine its properties, and if more than one gas is found to be present, determine the relative proportions of the gases.

4. Substance S is a mixture. Try, if you can, to separate the constituents, and find out their properties.

[1] "In the practical examination simple problems intended to test the candidate's powers of observation and description will be set on subjects selected from the schedule. Questions on practical mensuration will be set with one of the *papers*. The chemical exercises may include simple measurements of masses and volumes."

[2] P and Q were Lead and Cork.

R was Oxalic Acid.

S was Charcoal, Chalk, and Common Salt, or Sand, powdered Zinc, and Epsom Salts.

CAMBRIDGE LOCAL EXAMINATIONS

Junior Candidates

ELEMENTARY EXPERIMENTAL SCIENCE. Paper III

December 1900

2 Hours

No candidate may attempt more than 2 of the 4 questions. Full details as to observations made, together with a statement of the precautions taken to avoid error, are to be given.

A 1. Heat K, L, and M. Note what happens in each case, and compare your observations in the three cases.

A 2. Describe the action of heat and acids on the given carbonate N. Find the volume of carbonic acid given off on adding excess of nitric acid to a known weight of N, and calculate the number of c.c. given off by 1 gram. State the temperature and pressure at which you measure the gas.

B 1. Find the centre of gravity of the given figure by a graphical method.

B 2. Weigh a quantity (between 40 and 50 grams) of common salt, dissolve it in 250 c.c. of water, and determine the specific gravity of the solution.

JOINT SCHOLARSHIPS BOARD

Intermediate Scholarships Examination

ELEMENTARY EXPERIMENTAL SCIENCE

Practical Exercise

June 1900

½ Hour

You are provided with a white substance, and are required

(1) To describe its appearance and obvious properties.

(2) To determine its density by weighing it in air and in water.

When the superintendent has initialled your answer, make the experiment. Then describe exactly what you have done, and record your results.

JOINT SCHOLARSHIPS BOARD

Intermediate Scholarships Examination

ADVANCED CHEMISTRY

June 1900

½ Hour

With the materials provided you are required to fit up an apparatus such as would be suitable for determining the loss of weight incurred by chalk when it is treated with an acid. Mark the positions of the tubes by gumming rings of paper round them just above the cork, and indicate which is the exit tube. Put your name on a label, and also note on this the time taken in fitting up the apparatus.[1]

LONDON COUNTY COUNCIL

Intermediate County Scholarships

ELEMENTARY EXPERIMENTAL SCIENCE

Practical Work

July 1899

3 Hours

1. Measure the length and average diameter of the glass rod, and hence calculate its volume. Check your result by a direct determination of its volume by displacement of water in the burette. Calculate what length should be cut off to have a volume of 2 c.c., and cut off this length. Label it with your name and number.

2. Determine the specific gravity of the glass beads, using the balance and the specific gravity flask.

3. A is a mixture of a soluble and an insoluble substance. Obtain a dry sample of the insoluble substance, and find what volume of gas is given off from 2 grams of it by treatment with hydrochloric acid.

4. Prepare specimens of crystals from the solution B, and describe fully the changes which occur on heating them to a high temperature and treating the residue with hydrochloric acid.

[1] Each candidate was provided with a 4-oz. conical flask, a small straight chloride of calcium tube, and two corks which fit the flask well, gum paper, narrow glass, and rubber tubing, and narrow glass rod, all ⅛ in.

July 1900

3 Hours

1. Measure the diameter of one of the metal discs, and hence calculate the area of its face. Determine the volume of several of the discs by the aid of the beaker and the burette. Calculate the volume of a single disc, and hence the thickness of a disc. Check this result by a determination of the thickness directly, by measurement of the thickness of a number of the discs pressed together between the wooden blocks.

2. Determine by the aid of the balance and sinker the density of the substance given.

3. Determine the percentage by weight of carbon dioxide in the substance A.

4. Describe fully the changes which occur when B is heated and the residue obtained is treated with hydrochloric acid.

APPENDIX

Laboratory Fittings.—Apparatus and Chemicals

The Laboratory,[1] in which the student is to work, will probably be found by him ready fitted. If he has to fit up his own, the most important requirements are: –

A room of sufficient height, well lighted and ventilated.
A bench of some hard wood, such as teak, not easily destroyed.
Shelves for bottles above the bench, and on the walls.
Drawers and cupboards for apparatus, below the bench.
Gas for heating and lighting.
A water supply.
An earthenware or lead-lined sink, not easily choked by rubbish.
A draught-cupboard for substances evolving noxious fumes.

Plans for Laboratory fittings[2] may be obtained from:—

North of England School Furnishing Co., Darlington;
Hammer and Co., London;
John Heywood, Deansgate, Manchester;
The Science and Art Department's Drawings and Descriptions of Fittings for a Laboratory (Form No. 1013), Eyre and Spottiswoode; 1*s.* 6*d.*
Reynolds and Branson, Commercial Street, Leeds.

Choice of Apparatus.—When choosing apparatus to furnish a new laboratory, the funds available may be limited, but there are some directions in which a liberal initial expenditure would be wise economy, and other directions in which it is quite easy to postpone purchases.

The provision of identical apparatus, perhaps in sets of a dozen, means studied economy when a school has limited means. As to ways of saving, we need not waste money on expensive lecture-table

[1] Latin, *laborare*, to work.

[2] Illustrations of recently fitted laboratories will be found in the back numbers of the *Record of Technical and Secondary Education*. The standard book on laboratory planning is, however, "Technical School and College Buildings," by F. C. Roberts, F.S.A., demy 4to (Whittaker).

apparatus, to be used only once a year, and then only to show a single experiment. Hence we may have to omit the celebrated guinea-and-feather experiment, unless we can substitute for the time-honoured but extravagant apparatus some simpler contrivance of our own. A Hempel gas-burette, costing 18*s.*, and useful for an endless variety of experiments on the volumes of gases, is worth far more than the special apparatus, costing 25*s.*, and used only for the electrolysis of hydrochloric acid. Instead of specialised apparatus, it is better to buy a good supply of rubber corks, screw-clips, and T-tubes, and encourage the ingenuity of boys in building up more complicated apparatus from these simple elements.

But exact balances, one to every pair of boys, and accurate weights with each, although expensive, are indispensable.

Many old laboratories have become both untidy and inconvenient on account of the number of different patterns of bottles and of retort-stands which have accumulated by purchase at different dates. We therefore recommend a wise forethought in selecting carefully the best patterns of bottles, retort-stands, and boxes of weights, and the purchase of rather long stocks of these. (See also p. 204.)

APPARATUS FOR EACH BENCH

Each pair of students should have for their own use some such set of apparatus as the following :—

Retort-stand with two or three rings and one clamp.
One aspirator.
Two Bunsen gas-burners with rubber tubing.
Two cones for Bunsen burners.
T-piece with rubber tubing.
Tripod-stand.
Iron tray for sand.
Iron wire gauze.
Wicker basket for glass tubing and corks.
Duster.
Three flasks : 16 oz. and two 8 oz. conical flasks.
Test-tube stand.
Six test-tubes, 5 × ⅝ in.
Test-glass.
Glass tumbler.
Crucible tongs.
Two files, triangular and rat-tail.

Small pestle and mortar; Berlin, glazed, 2 in. diameter.
Two porcelain evaporating basins.
Two filter funnels.
Thistle funnel.
Two watch-glasses.
Two glass rods.
Two pipe-clay triangles.
Mouth blow-pipe.
Narrow rubber tubing—two or three small pieces.
Mohr's spring clip.
Deflagrating spoon.
Broad burner.

APPARATUS FOR GENERAL USE

In addition, the following apparatus will be frequently wanted, and must be close at hand, and in sufficient quantity:—

Flasks, flat-bottoms, German glass; various sizes (24, 16, 8, 4, 2 oz.).
Flasks, conical; 8 and 4 oz.
Flasks, round-bottoms, hard Bohemian glass; 16 and 8 oz.
Bottles, wide-mouthed; 12 oz. (preparation of hydrogen, etc.).
Beakers; various sizes down to 1 oz.
Beakers, lipped; 14 and 8 oz.
Test-tubes; 3 × $\frac{1}{2}$, 5 × $\frac{5}{8}$, 6 × 1 in.
Hard glass test-tubes; 6 × $\frac{3}{4}$ in.
Hard glass reduction tubes, 1 bulb.
Arsenic bulb tubes.
Specimen tubes; 2 × $\frac{1}{2}$ and 2 × $\frac{3}{4}$ in.
Glass retorts, stoppered; 8 oz.
Sample bottles; 8 and 4 oz.
Winchester pint bottles.
Bell-jars.
Soft glass tubing; $\frac{1}{4}$ and $\frac{3}{16}$ in.
Hard glass tubing; $\frac{1}{2}$ in. internal diameter.
Barometer tubing; 36 in.
Glass evaporating basins; 3 in. diameter.
Glass crystallising dishes; 3 in. diameter.
Watch-glasses; 3, 2$\frac{1}{2}$, 2 in.
Calcium chloride bulb tubes, 6 in.
,, ,, U-tubes.
Two large funnels; 8 in. diameter.

Separating funnels ; 100 c.c.
Lamp-glasses.
Test-glasses ; 4 oz.
Measuring flasks, stoppered : 1,000, 250, and 100 c.c.
Measuring jars ; 500, 200, and 100 c.c.
Pipettes ; 50 c.c., 10 c.c., and 100 grains.
Burettes, Mohr's ; 50 c.c., with clip and jet.
Gas jars ; 8 × 2, and 6 × 1¼ in.
Round ground glass plates ; 3 in. diameter.
Glass T-tubes.
Wood scales[1] ; metre, ½ metre, and 30 cm. + 1 ft.
Filter paper ; cut in circles, 4½, 3½, 2½ in. diameter.
,, ,, in sheets.
Squared paper (see Vol. I, p. 188).
Porcelain crucibles ; 1½ and 1¼ in. diameter.
Wedgewood ware pestles and mortars ; 5 in., and one large 8 in.
Iron pestle and mortar.
Agate pestle and mortar.
Thermometers ; Centigrade, reading to 200° and 110°
,, Fahrenheit, reading to 240°.
Pneumatic troughs.
Beehive stands.
Porcelain boats ; ½ in. and ⅜ in. wide.
Black rubber tubing ; ¼ in. and $\frac{3}{16}$ in. diameter.
Red rubber tubing ; $\frac{5}{16}$ in. diameter.
Pump tubing.
Corks ; 1, ⅞, ¾, ⅝, ½, ⅜, ¼ in. diameter.
Bungs : 1⅝ and 1 in. diameter.
Rubber stoppers : 1 hole ; ⅞, ¾, ⅝, ½, ⅜ in. diameter.
,, ,, 2 holes ; 1⅛, 1, ⅞, ¾ in. diameter.
,, ,, solid ; ⅞, ¾, ⅝, ½ in. diameter.
Blocks of wood.
Cedar splints.
Balances ; to carry 250 grams and 100 grams, one for every two students.
Sets of weights ; up to 100 grams and down to 1 centigram.
Salter's spring balance ; to carry 1,000 grams.
Béranger balance (for rapid and rough weighing).
Set of weights ; up to 1,000 grams, and down to 1 decigram.
Spare decimal of gram weights.
Set of grain weights.
Ramsay burners.

[1] From J. Raybone and Sons, Whitmore Street, Birmingham.

Fletcher's radial burners.
Iron spoons.
Brass and zinc cans for calorimetry.
Screw clips.
Labels: blank.
Labels: printed and varnished.
Small retort-stands.
Cylindrical cans, for boiling water; quart size.
Platinum wire and foil.
Emery paper.
Battery clamps.
Cubes and blocks of hard wood.
Copper, brass, and iron cylinders; $1\frac{1}{2} \times \frac{5}{8}$ in.
Copper, lead, and iron balls.
Small horse-shoe magnets.
Balloons, goldbeaters' skin; 18 in. circumference.
Pliers.
Wooden trays (1 doz.) for mercury experiments; for carrying apparatus about.
Specific gravity bottles; 25 c.c.

APPARATUS FOR OCCASIONAL USE

Grove's battery, bichromate battery, or secondary cells.
Small induction coil.
Kipp's apparatus (two, for carbonic acid gas and hydrogen).
Hempel gas burettes.
Combustion furnace.
Muffle furnace. (Fletcher's, 23H.)
Foot-bellows and blow-pipe.
Brass syringe, exhausting and condensing.
Air pump, Tate's.
Water vacuum pump.
Drying oven, 6 in. square.
Water bath.
Cryophorus.
Wet and dry bulb thermometers.
Eudiometers (U-shaped).
Liebig's condenser.
Balance turning with 0·1 milligram.
Dessiccators.
Spectroscope.

CHEMICALS FOR INDIVIDUAL USE

If qualitative analysis is not being done in the laboratory, only a very few bottles need be provided at each place, say—

Sulphuric	acid	Ammonia
Nitric	,,	Caustic soda (rubber stopper)
Hydrochloric	,,	Lime water.

CHEMICALS FOR GENERAL USE

Acids—Acetic.
Hydrochloric, commercial and pure.
Nitric, fuming, commercial and pure.
Oxalic, pure.
Sulphuric, commercial and pure.
Tartaric.
Alcohol—Methylated.[1]
Amylic, pure.
Alum—Ammonia.
Chrome.
Iron.
Potash.
Aluminium, sheet.
Ammonia solution.
Ammonium—Acetate.
Carbonate.
Chloride.
Nitrate.
Antimony, powder.
Asbestos.
Barium Chloride, crystallised.
Benzene.
Bismuth Oxide.
Borax.
Calcium—Carbonate (chalk), B.P.
Chloride, fused.
Hypochloride, commercial and in cubes.
Oxide (quicklime), pure.
Carbon Bisulphide.

[1] The [illegible] kind of unadulterated methylated spirit may be obtained from the Vauxhall Methylating Company, Liverpool, by arrangement with the Inland Revenue Department.

Castile soap, white.
Charcoal, picked sticks.
Chloroform, pure, B.P.
Chromium Oxide, green.
Cobalt Oxide.
Copper—Metal, pure foil, turnings, and wire.
Carbonate, pure.
Nitrate.
Oxide, black, pure.
Sulphate, pure and commercial.
Dutch Metal.
Ether, methylated.
Fusible Metal.
Fusion Mixture.
Glycerine.
Gold Leaf.
Gum Arabic.
Iodine.
Iron—Metal, filings, pure wire, and reduced by hydrogen.
Ammonium sulphate.
Oxide, magnetic and sesqui-.
Sulphate, pure.
Lead—Metal, foil and granular.
Carbonate.
Nitrate.
Red.
Oxide (litharge).
Peroxide.
Litmus—Pure granular.
Books, red and blue
Logwood.
Magenta.
Magnesium—Metal, ribbon, and wire.
Carbonate, light.
Nitrate.
Oxide.
Sulphate, crystallised.
Manganese, Dioxide (free from carbon).
Methyl orange.
Mercury—Metal.
Iodide (red).
Nitrate.
Oxide (red).

Olive oil.
Paraffin, solid white.
Phenol phthalein.
Phosphorus, yellow and red.
Plaster of Paris.
Platinum—Wire and foil.
Potassium—Metal.
Bichromate.
Carbonate, pure, dry.
Chlorate.
Chloride, pure.
Cyanide.
Ferricyanide.
Ferrocyanide.
Hydrate, sticks, pure.
Nitrate.
Nitrite, sticks, pure.
Permanganate.
Sulphate.
Pyrogallic Acid.
Silver—Metal, leaf.
Nitrate.
Sodium—Metal.
Bicarbonate.
Carbonate, anhydrous, light, pure.
Chloride, pure, and fused.
Formate, commercial.
Hydrate, lump and sticks.
Nitrate.
Potassium Tartrate (Rochelle Salt).
Sulphate.
Thiosulphate.
Solder.
Starch.
Sugar, cane.
Sulphur, roll and flowers.
Test-papers, litmus.
Tin—Metal, granulated, and foil (pure)
Oxide (white).
Toluene.
Turpentine.
Water, distilled.
Zinc—Metal, granulated, and dust.

Zinc—Metal, foil, pure.
Oxide.
Sulphate, crystallised, pure.

Minerals—Azurite.
Calc-spar.
Cassiterite.
Chalk.
Cinnabar.
Copper—Native.
Pyrites.
Cuprite.
Flint.
Galena.
Gypsum.
Hæmatite, red.
Iron Pyrites
Marble, white.
Quartz, crystal.
,, with gold.
Rock Salt, clear white.
Sand, yellow and white.
Silver glance.

NOTES ON APPARATUS

The Laboratory should be arranged so that every one who uses the room may know where to find everything he wants and where to put it away again. There should be numbered places for everything. Numbers on boxes and bottles, and printed lists on the walls, will soon repay the initial trouble of preparing them in the time they save. It should also be possible to put apparatus aside. A dozen or more trays measuring 27 in. × 15 in., and a gross of empty wooden boxes, 12 in. × 8 in. × 4 in., such as those in which sweetmeats are packed, will greatly help. These should be labelled with letters or numbers to indicate their place, and the nature of the contents should be clearly marked on the outside of each.

Labels should be freely used, and all unlabelled bottles washed out. Paper labels on bottles may be protected by a coating of paraffin wax, put on when the gum is dry and the bottle hot; or the label may be first sized and then varnished; or "Primus indestructible labels" may be obtained ready varnished. For numbering, the "auctioneers' lotting labels" supplied by Wilson and Son, 68 Cloudesley Road, London, are most useful.

But most bottles should have some label in the substance of the glass. These are made in two forms. Enamel labels are melted on to the glass. They last well, but white enamel is not always distinct when white crystals are inside the bottle. Sand-blast labels are made by roughening the surface of the glass with sand driven in a blast of air. They are apt to be indistinct when wet, if not made with a really rough surface. They can be made more distinct by rubbing over with blacklead or chalk. Sand-blast labels cost only about half as much as enamel labels. But, whichever sort of label is adopted, every bottle should bear some letter or number to show its place as well as its name.

Retort Stands should be firm and substantial. It is worth while to pay five shillings or more to get good ones. The sliding collars for the rings should be slotted at the sides, so as to be moved easily; and good clamps should be chosen, with well-made screw threads.

Balances should turn with much less than a centigram. The Dutch pattern balances, introduced some ten years or more ago, continue to give great satisfaction, and since then the quality has improved and the price has gone down. Balances with steel knife-edges, if protected from the air of the laboratory, when not in use, within a good cupboard or under a glass case, will remain serviceable for five or six years. Those with agate edges set in steel are more durable, but more expensive. It is desirable that they should be kept and used on shelves at the side of the laboratory, or in a separate balance-room, where they need not be moved.

Boxes of Weights.—Many sets recently placed on the market are pretentious but inaccurate, the errors of the largest weights being greater than the smallest weights. Weights less than a centigram are needless, and cause trouble to beginners. More exact weighing can always be done by watching the pointer if the weights are accurate. The set we like extends from 100 grams to 0·01 gram, and has a separate space, no more no less, for each weight. It was put up by Messrs. Oertling at 18*s.* 6*d.* per box, to answer the test that no single weight nor combination of weights should have an error greater than 0·005 gram.

Messrs. J. J. Griffin and Sons now advertise a set (5651), from 100 grams down to 0·001, limit of error 3 milligrams, price 8*s.* 6*d.*, and this corresponds to our specification, except in the presence of weights less than 1 centigram, whose absence we consider desirable, and in putting two weights of the same value into one hole, which makes the boxes less easy to keep in order with classes of beginners.

Glass Bottles, of similar general appearance, but of very different real quality, may be obtained at astonishingly different prices. Hand-made bottles have a strength of substance and lustre of surface wanting in moulded bottles. Flat-topped stoppers keep dust out of the neck; upright stoppers are more easily removed with the back of the fingers. English stoppers are better fitted than Continental ones. 8 oz. bottles are a convenient size for solutions in individual use, 20 oz. for solutions in general use, 80 oz., or Winchester quarts, for larger store. Dry solids are conveniently kept in 6 pint stoneware jars, with 16 oz. and 6 oz. bottles for smaller sizes.

To begin with, 40 stoneware jars, 20 Winchester quarts, 10 dozen smaller bottles for solutions, and 10 dozen bottles for solids, will not be too many to order if the laboratory is to be arranged for some years to come.

Deflagrating Spoons are cleaned by holding the cups in the Bunsen flame for a few seconds.

The Grove's Battery.—When the cell is taken to pieces, the nitric acid may be kept to use again. The dilute sulphuric acid and zinc

sulphate from the outer cell must generally be thrown away; the nitric acid may be used again, but only for the same purpose. The porous cells should be soaked in fresh water for a day or two, with one or two changes, until the water tastes quite sweet. Then allow them to dry. The zinc plates should also be soaked in water. Before fitting up the Grove cell all ends of wires and clamps must be brightened with a file, or with sand-paper, to ensure a good metallic contact.

The zinc plates may be cleaned by scraping and washing (if greasy wash with caustic soda). Dip them into dilute hydrochloric acid, and wash again. Rub some mercury over the surface of the plates. Dip them again in acid and then pour more mercury over them. Repeat this until the zinc will stand in the acid without effervescing. The mercury may be used again for the same purpose, but it should be kept separate from the general stock.

To charge the battery, fill the porous cells with fuming nitric acid. Fill the outer cells two-thirds full with a mixture of one part of sulphuric acid and eight parts of water.

Ruhmkorff Coil—Bichromate Cell.—For a full account of a Ruhmkorff coil, the student must refer to some book on electricity, such as Silvanus Thompson's *Electricity and Magnetism*. For the present he may understand it as a machine for converting a steady continuous current of electricity into an intermittent series of violent sparks. The coil may be used in connection with the Grove's battery, but as this takes some time to set up it is more convenient to use a bichromate battery which can be thrown in or out of action in a moment by raising or lowering the zinc plates. The zinc plates ought to be amalgamated like those of the Grove's battery. The bichromate battery is charged with a mixture in the proportion of

Potassium bichromate	13 gm.
Water	140 c.c.
Sulphuric acid	24 c.c.

This mixture is of a clear orange colour when first prepared, but gradually changes to dark green and deposits crystals of chrome alum in the cell.

Kipp's Apparatus (Fig. 67) is the best for supplying a continuous current of hydrogen unmixed with air. The bulbs B and C are of one piece. Into these is fitted the funnel A. Removing A and F, zinc rods can be slipped into B. The glass collar D is pierced with holes allowing liquid to drain down into C, but stopping any solid material. The dilute sulphuric acid should be cooled before putting it in, as the thick glass is apt to crack. But hydrochloric acid will do at least as well, and saves this risk. To fill the apparatus, open the tap F. Pour acid into

A until C is full. Air is driven out through F. Close F. Pour in more acid until A is two-thirds full. Acid cannot enter B because B is full of air. Open the tap F. Air escapes under the pressure of the acid column in A. Acid enters B, and effervescence begins. Close F. The effervescence continues, but the hydrogen, being unable to escape, accumulates in B. The acid is driven back into C, and up into A. The effervescence stops as soon as the acid has drained off the zinc. The stopper E is useful when washing out the apparatus. Sulphuretted hydrogen, hydrochloric acid, and carbonic acid gas may be prepared in the same apparatus.

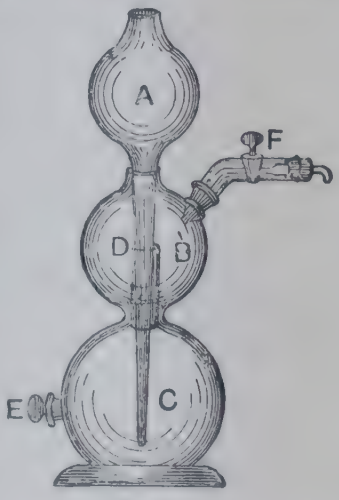

Fig. 67.

Management of Mercury.—Mercury is difficult to manage, on account of its weight and the ease with which it picks up impurities. Stoneware mortars are better for holding it than glass beakers, which easily crack when put down on the bench.

All work with mercury should be done over a tray, or on a table with a groove round the edge.

Dust on the surface may be removed by letting the mercury run through a dry filter-paper in which a few holes have been pricked. Or the mercury may be washed in a current of water, and then passed through a separating funnel. A little moisture on the surface may be removed with blotting-paper.

Other metals easily dissolve in mercury; and a very little tin or zinc will spoil its fluidity. They are best removed by shaking the mercury in a bottle with a little nitric acid or nitrate of mercury, when zinc and tin will dissolve in preference to the mercury.

Bunsen Burners occasionally get blocked by molten substances falling into the inner tube. They may be cleaned by washing with water, or by picking it out with a pin. After washing, blow through the burner, to remove any drop of water which may clog the exit. The revolving tube or ring ought to be frequently turned, so as to keep it in order; otherwise they get jammed, and it is no longer possible to change quickly from yellow to blue flames.

ANSWERS

Examples I

The Densities of Gases. (Page 117.)

1. 1·28 gram.
2. 1·$\dot{3}$ gram.
3. 775·19 c c.
4. 1·215 oz.
5. 124·7 grams.
6. 1011·85 grams.
7. 1·56 gram.
8. 60 grams.
9. 0·68 gram.

Examples II

Boyle's Law. (Page 127.)

1. 735·15 or $735\frac{15}{101}$ c.c.
2. 1500 c.c.
3. 742·57 or $742\frac{58}{101}$ c.c.
4. 288 cu. ins.
5. 1200 lbs. per sq. in.
6. 40 mm.
7. 43·48 or $43\frac{11}{23}$ c.c.
8. 4 ins.
9. 60 ins.
10. 250 mm.
11. 0·098.
12. 1 000.

Examples III

Charles's Law. (Page 137.)

1. (1) − 272°·97 C.
   (2) − 271°·51 C.
2. (1) $\frac{1}{272\cdot8}$ or 0·003665.
   (2) $\frac{1}{270\cdot13}$ or 0·00370.
3. 954·54 c.c.
4. 1076·92 c.c.
5. 1·17 grams.
6. 380·0082 grams.

Examples IV

Charles's and Boyle's Law. (Page 138.)

1. (*a*) 79·2 c.c.
   (*b*) 82·11 c.c.
   (*c*) 109·87 c.c.
2. (*a*) 556·78 c.c.
   (*b*) 560·44 c.c.
   (*c*) 476·19 c.c.
3. (*a*) 322°·67.
   (*b*) 162°.
   (*c*) 46°.
4. (*a*) 709·8 mm.
   (*b*) 599·04 mm.
   (*c*) 907·2 mm.

Examples V

The Densities of Gases. (Page 138.)

1. 1·44 grams.
2. 0·5050 litre.
3. 1646·09 c.c
4. 0·5865 gram.
5. 103·9 litres.
6. 0·0264 gram.
7. 4095° C.
8. 319° C.

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